HEATH

Microscale Laboratory Manual

CHEMISTRY

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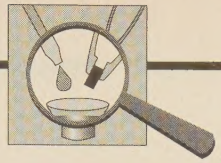

Heath

CHEMISTRY

Microscale Laboratory Manual

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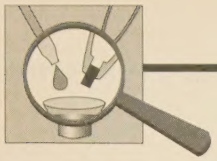

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Experiment 1	*Measurement & the SI System*	1
Experiment 2	*Conversion Factors*	5
Experiment 3	*Classification of Mixtures*	11
Experiment 4	*Separation of Mixtures*	15
Experiment 5	*Recognizing Chemical Reactions*	21
Experiment 6	*Interpreting Chemical Reactions*	25
Experiment 7	*Mass Relationships in Compounds*	29
Experiment 8	*An Investigation of the Mole*	35
Experiment 9	*Identifying an Unknown Oxide*	41
Experiment 10	*Synthesis of Manganese Chloride*	45
Experiment 11	*Gas Laws I*	49
Experiment 12	*Gas Laws II: The Universal Gas Constant*	57
Experiment 13	*Classifying Elements*	61
Experiment 14	*Activity Series I: Metals*	65
Experiment 15	*Activity Series II: Nonmetals*	69
Experiment 16	*Dyes and Dyeing*	73
Experiment 17	*Vapor Pressure of Water*	77
Experiment 18	*Double Replacement Reactions in Solution*	81
Experiment 19	*Calorimetry*	87
Experiment 20	*Investigating Reaction Rates*	93
Experiment 21	*Testing Vinegar Concentration*	99
Experiment 22	*Redox Reactions*	105
Experiment 23	*Testing for Water Hardness*	111
Experiment 24	*Separating Liquids by Distillation*	117

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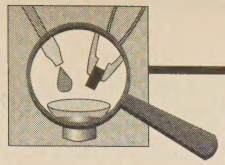


To the Student

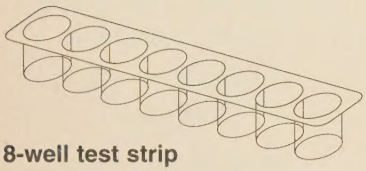
A microscale laboratory program follows the philosophy that *less is more.* For an experiment to qualify as microscale or small-scale, it must use smaller quantities of chemicals than are typically used in conventional macroscale experiments. One advantage of smaller-scale labs is a reduction in the cost of chemicals. Another is a lower risk of exposure to hazardous chemicals. A third advantage is a reduction in the amount of waste. The experiments in the *Heath Chemistry Microscale Laboratory Manual* are designed to have minimal environmental impact. Each experiment contains a section called *Cleaning Up* with instructions for the safe disposal of unused reagents and products.

An added advantage of small-scale labs is the reduced time needed to complete an experiment. The extra time can be used to complete more labs in one year or to run additional trials of a single lab, which parallels the approach of chemists in research laboratories. Faster procedures also provide more time for communication—an essential step in scientific discovery. For results to be meaningful, they must be reproducible by the initial investigator and by others. During these experiments, you will often be encouraged to consult with others, to compare results, and to draw conclusions based on those comparisons.

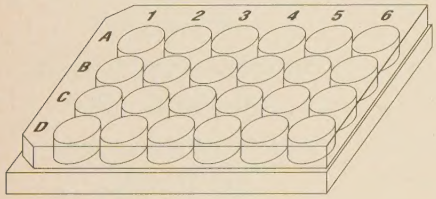
Chemistry is a process. It is concerned with investigating the behavior of matter and with interpreting the results of those investigations. Your goal when doing an experiment is not to find "the right answer." Your goal is to report what you observe. Assuming you do not make errors while recording the data, what you observe is always *the right answer.* Differences in reported results among teams can be traced back to a difference in the procedure followed. Much can be learned from exploring such differences.

There is one disadvantage to working with smaller quantities. Slight errors in measurement are magnified. For example, one or two drops can represent a significant percent of the total volume being measured. You must work with increased precision to minimize errors. However, if you answer the Prelab questions, read the introduction and procedure before beginning, work in unison with your partner, and are attentive to detail, you can successfully complete all experiments.

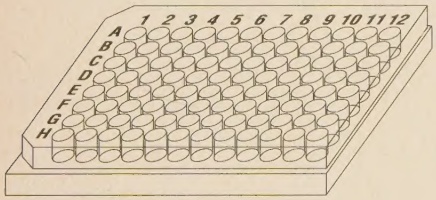
Microscale Equipment



8-well test strip



24-well plate



96-well plate

microtip pipet

thin-stem pipet

Caution Do not use method 2 when the pipet contains acids, bases, or other corrosive liquids.

Smaller quantities of chemicals require either reduced-size standard equipment (such as 30 mL beakers or 10-mL graduated cylinders) or different categories of equipment—micropipets and test strips or plates.

Test strips or plates serve as substitutes for the beakers and test tubes used in larger-scale labs. The strips contain 8 or 12 wells, each with a capacity of about 0.3 mL. The wells in a 96-well test plate have the same capacity as those in a test strip. The wells in a 24-well plate have about 10 times the capacity of a well in a 96-well plate—about 3 mL. As you will discover in Experiment 3 when you attempt to rinse out your test strip, plastic test strips have some unexpected features.

Plastic pipets (sometimes referred to as Beral pipets after one of the first companies to manufacture them) come in many sizes and shapes. Most of the time you will use thin-stem pipets. Sometimes you will use microtip pipets. The main difference between these two types is the size of the opening on the pipet tip. You will investigate this difference further in Experiment 1.

The key to successful pipet handling is to develop a technique that works for you and to use it in every experiment requiring pipets, whether thin-stem or microtip. In the early experiments, you will have an opportunity to establish your technique. Here are a few tips to help you get started.

1. Hold the pipet at a consistent angle. When adding drops to a narrow container, it is best to hold the pipet vertical to the container. You do not want drops to adhere to the sides of the container.
2. Apply a smooth, steady pressure on the bulb of the pipet. Try to avoid squeezing the bulb with your fingertips ("pinching" the bulb). Pinching tends to release multiple drops at the same time.
3. Start with a pipet that is at least half full. It is difficult to maintain a steady pressure on the bulb when the pipet is almost empty. When you are repeating a procedure two or three times, check the amount of material in the pipet before each trial.
4. The bulb and stem of a pipet do not need to be completely full of liquid for a pipet to qualify as filled. A *filled* pipet actually contains about one-quarter to one-third empty space.
5. Small amounts of liquid can get caught in the stem of the pipet, which can lead to the expulsion of air bubbles from the tip of the pipet. To avoid bubbles, store the pipet in a bulb-down position when not in use.

   There are two techniques for dislodging bubbles. (1) Just before use, tap the bulb gently on your work surface to dislodge the liquid in the stem so that it flows back into the bulb. (2) Just before use, with the bulb still pointing down, gently expel the liquid in the stem into a folded paper towel held over the sink. The second method works only when your results are based on counting drops, not when you are measuring the amount of liquid in the pipet before and after the procedure.

Safety in the Laboratory

In the laboratory, you will reinforce and extend the concepts you learn in *Heath Chemistry*. You need to work with caution to ensure an accident-free environment for you and your classmates. The information below describes basic rules for safe laboratory practice. Your instructor will provide additional guidelines specific to your laboratory setting. If you adhere to these guidelines, your laboratory experience should be both safe and rewarding.

1. Do only those experiments assigned by your teacher. Never do an experiment without your teacher's supervision.
2. Locate and learn to use all safety equipment.
3. Follow instructions concerning the wearing of safety goggles, aprons, and gloves. (Everyone will be wearing the same gear, so you will all look equally in fashion.)
4. Regular glasses are preferable to contact lenses for laboratory work. Vapor from chemicals can be trapped between contact lenses and your eyes. Safety goggles do not provide protection against such vapor. If you do wear contact lenses during lab, mark your goggles so that others will know you are wearing lenses in case of an emergency.
5. Do not wear clothing with loose-fitting sleeves, scarves, dangling jewelry, or open-toed shoes to lab. Tie back long hair.
6. Treat all reagents with due respect, even those that carry no special warning. Unless instructed otherwise, do not handle chemicals directly. Never taste a reagent. If instructed to smell a chemical, keep your nose at least 20 cm away from the container. Wave your hand gently over the mouth of the container to waft the vapor toward your nose.
7. Vapor from chemicals can be absorbed by other items in the laboratory. Do not eat, drink, chew gum, or apply cosmetics while in the laboratory. Do not store food or beverages in the lab area.
8. Chemicals can be absorbed through the skin. Avoid touching sensitive areas of your skin, such as your face, during lab. If a chemical gets in your eyes, use an eyewash fountain *immediately*. If reagents spill or splash on your skin, rinse the affected part immediately with a continuous stream of cool tap water. If the reagent is corrosive, flush the affected area for at least 15 minutes. If corrosive chemicals are splashed on an item of apparel, remove the item while under a safety shower.
9. Do not work with flammable chemicals if there are open flames in the laboratory. Keep flammable chemicals away from sources of spark and heated surfaces (on hot plates or sand baths).
10. Use a clamp, sturdy tongs, or a heat-resistant mitt to handle hot objects. Never place a hot object directly on a balance pan. Never leave heat sources, such as sand baths, unattended.
11. Keep electrical equipment away from sources of water. Dry your hands before touching electrical equipment. Notify your teacher if you notice frayed cords, loose connections, or broken wires. Make sure cords do not dangle from work tables. Your teacher will connect and disconnect sand baths and hot plates. Do not change the settings on sand bath controllers. In case of an emergency (such as a spill near an electrical device), disconnect an appliance from an outlet by pulling on the plug, not the cord.
12. Horseplay or practical jokes are not appropriate in a chemistry laboratory. Equipment should be used *only* for designated purposes. It is your responsibility to avoid immature behavior and it is your right to expect the same of others.
13. Keep your work area uncluttered. Store books other than your lab manual and items such as backpacks in designated storage areas. Keep apparatus and reagents away from the edge of your lab bench.
14. Report all accidents—burns, spills, cuts—to your teacher immediately, no matter how minor the incident may appear. Do not handle broken, cracked, or chipped glassware.
15. Do not return an unused reagent directly to a stock bottle. Follow the specific instructions for disposal of reagents in the *Cleaning Up* sections. Unless instructed otherwise, clean test plates, well strips, glassware, and your work area at the conclusion of each lab period. Wash your hands thoroughly with soap and water before leaving the lab.

Safety Logos

Study these safety symbols. They are used as visual tags for the *Cautions* that appear throughout this manual. Review the entire procedure, especially the *Cautions,* before you begin work on an experiment. If you have questions, ask your instructor during Prelab.

Wear safety goggles *whenever* you are working in the laboratory.

Wear a laboratory apron to prevent damage from corrosive chemicals, hot materials, or reagents that can stain clothing.

Wear gloves as protection against corrosive or irritating reagents.

This symbol indicates a corrosive or irritating reagent. A corrosive reagent can damage living tissue. An irritant can cause inflammation.

A face shield provides additional protection from chemical splashes. Reagents requiring the use of a face shield will be handled by your instructor.

This symbol indicates a flammable reagent or an apparatus that can produce burns—sand baths, hot plates, or laboratory burners.

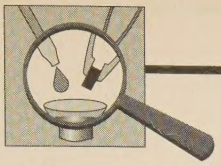
This symbol indicates electrical equipment such as sand baths or hot plates.

This symbol indicates scissors or other sharp objects.

This symbol indicates a toxic reagent. Toxins must be treated with considerable caution. They can enter the body by various routes—through the digestive and respiratory systems, through cuts in the skin, and sometimes even through intact skin.

The fume hood is used for procedures involving reagents with toxic fumes.

This symbol alerts you to safety procedures that are specific to a particular experiment.



Experiment 1

Measurement & the SI System

Scientific observations often include measurements of volume, mass, length, and temperature. With measuring instruments such as rulers or graduated cylinders, readings may usually be estimated to one more decimal place than the smallest scale division. Thus, a ruler that is divided into millimeters allows you to estimate to the nearest 0.1 mm. A few devices, such as electronic balances, give a direct reading—you cannot estimate between scale divisions. Sometimes a piece of equipment will have scale markings that are only approximate and are not intended to be used for precise work at all. This is true of the beakers and Erlenmeyer flasks that you will use in this course.

Although the activities in this experiment are numbered, they can be done in any order. Your teacher may assign different groups to begin with different parts of the procedure. If not, organize your work to adjust for availability of equipment.

OBJECTIVES

1. ***Compare*** the precision of various types of liquid-measuring containers.
2. ***Determine*** average dimensions for coins.
3. ***Measure*** and ***compare*** the masses of pre-1982 and post-1982 pennies.
4. ***Determine*** the effect of salt on the temperature of an ice–water mixture.

▶ Materials

Apparatus

centigram or milligram balance
metric ruler
50-mL beaker (with calibrations)
100-mL beaker
25-mL graduated cylinder
100-mL graduated cylinder
thin-stem pipet
small cup (optional)
thermometer, –20°C to 110°C
several pennies,
pre-1982 and post-1982
nickel coin
safety goggles
laboratory apron

Reagents

tap water
ice, small cubes or crushed
rock salt

▶ Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–4 on the Report Sheet.

▶ Procedure

Part 1: Reading Volumes

Put on your laboratory apron and safety goggles.

There are five different graduated cylinders, each labeled and each containing some quantity of liquid to which food coloring has been added to make the volume easier to read. Record the capacity of each cylinder and the volume of liquid in Table 1 on the Report Sheet. Remember to include the units (mL).

Part 2: Comparing Measuring Containers

Use tap water to fill a 50-mL beaker to the 20-mL mark. Use a thin-stem pipet to adjust the level until the bottom of the meniscus is lined up with the 20-mL line.

Pour the water from the beaker into a 25-mL graduated cylinder. Read the volume according to the 25-mL graduate and record it in Table 2. Finally, transfer the water from the 25-mL graduate to a 100-mL graduate. Read the volume again and record it in Table 2.

Part 3: Measuring Coins

Use a metric ruler to determine the diameter of a penny and a nickel. Report the diameters in both centimeters (cm) and in millimeters (mm). To the nearest whole number, determine the number of pennies it takes to make a stack that is 1 cm tall. Find the mass of this number of pennies. Record these measurements in Table 3.

Part 4: Comparing Pennies.

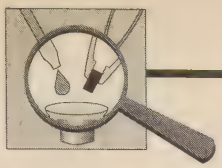
Determine the masses of two pennies. Make sure one is dated before 1982, and the other is dated after 1982. Record the masses in Table 4.

Part 5: Effect of Salt on the Temperature of an Ice–Water Mixture

Half-fill a 100-mL beaker with a mixture of ice and water; about half of the mixture should be ice. Place a thermometer in the ice–water mixture and hold it there until the temperature is no longer changing (usually 15–20 seconds). Read and record the temperature in Table 5.

Add one level spoonful of rock salt. Do not stir the solution. When the temperature is again no longer changing, read and record the final temperature of the ice–saltwater mixture.

▶ ***Cleaning Up*** Pour the salt-water–ice mixture down the drain. Rinse the beaker and thermometer. Return all equipment—including glassware, coins, and rulers—to their proper locations.



Experiment 1 Report Sheet

Measurement & the SI System

Name _____

Class _____

Date _____

▶ Prelab Questions

1. If you are using a graduated cylinder whose smallest division is 0.1 mL, to what degree of precision should you report liquid volumes?

2. A stack of five 3.5-inch floppy disks is 1.6 cm tall.

a. What is the average thickness of one disk?

b. To the nearest whole number, how many disks will be in a stack 10 cm tall?

3. In Celsius degrees:

a. What is the boiling point of water?

b. What is the freezing point of water?

c. What is normal human body temperature?

4. What is a meniscus? What role does it play in the correct reading of liquid volumes?

_____

_____

▶ Data and Observations

Table 1

Graduated Cylinder	Capacity	Volume of Liquid
1		
2		
3		
4		
5		

Table 2

Volume in 25-mL graduate:

Volume in 100-mL graduate:

Table 3

Diameter of penny: _____ cm _____ mm

Diameter of nickel: _____ cm _____ mm

Pennies in 1-cm stack:

Mass of 1-cm stack:

Table 4

Mass of pre-1982 penny:

Mass of post-1982 penny:

Table 5

Temperature of ice–water mixture:

Temperature of ice–salt–water mixture:

▶ Analysis and Conclusions

1. Did you get the same volume reading from each container in Part 2? If not, what explanation can you offer

   a. for an apparent decrease in volume?

   _____

   b. for an apparent increase in volume?

   _____

2. Which container in Part 2 gave you the most precise reading of the actual volume of water it contained? Explain.

   _____

3. The beaker you used in Part 2 probably carries the notation "±5%." What do you think this means?

   _____

   _____

4. Calculate the average thickness of a single penny.

5. Explain how you could estimate the mass of a large number of pennies, using only the data in Table 3.

   _____

   _____

6. Determine the number of pennies that could be laid edge-to-edge the full length of a meterstick. Repeat the calculation for nickels.

   _____ pennies

   _____ nickels

7. Compare your results from Part 4 with those of other teams. How do the masses of pennies minted before 1982 compare with the masses of pennies minted after 1982? Try to explain the difference.

   _____

   _____

8. Suggest two reasons why different teams did not report identical answers for Part 4 *other than* errors in reading the balance.

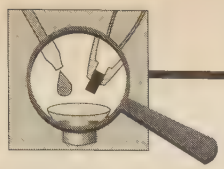
   _____

   _____

9. What effect does salt have on the temperature of an ice–water mixture? Did other groups observe the same effect?

   _____

   _____



Experiment 2

Conversion Factors

In this experiment you will be introduced to the derivation and use of conversion factors. As the name implies, conversion factors are a means to change from one unit of measurement to another. Some of these factors are familiar from everyday use: seven days is one week, twelve inches equals one foot, and so on. In each case, the same quantity is represented in two different ways. While seven days is the same length of time as one week, no one would suggest that the numbers themselves, seven and one, are equal—that would be absurd! Thus,

7 days = 1 week but $7 \neq 1$

12 inches = 1 foot but $12 \neq 1$

100 centimeters = 1 meter but $100 \neq 1$

In each example, quantities are equal, but numbers are not, and this distinction between numbers and quantities is critical. You must include the units in all reported measurements; without the units, all you have is a number—no quantity has been represented.

OBJECTIVES

1. ***Determine*** the mass of several known volumes of water.
2. ***Determine*** the volume of specific numbers of drops.
3. ***Graph*** the mass–volume data for water.
4. ***Derive*** conversion factors for changing one unit of measurement to another.

▶ Materials

Apparatus

milligram or centigram balance
thin-stem pipet
30-mL beaker (for distilled water)
10-mL graduated cylinder
safety goggles

Reagents

distilled water

▶ Prelab

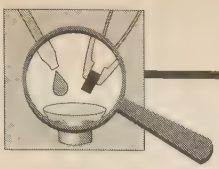
1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–3 on the Report Sheet.

Procedure

1. Put on your safety goggles.
2. Find the mass of a clean, dry 10-mL graduated cylinder to the nearest milligram or centigram, depending on the balance. If the only cylinder available is wet inside, you should dry it completely before you begin. Record the mass in Table 1 on the Report Sheet.
3. Use a beaker of distilled water and thin-stem pipet to add precisely 1.50 mL of water to your graduated cylinder. The low point of the curved liquid surface (the meniscus) should just line up with the 1.50-mL mark on the scale. Read at eye level to be certain that you gauge the volume correctly. Hold the pipet vertically so that the drops will fall freely to the bottom of the cylinder. If drops are left on the side walls of the graduate, your results will be inaccurate.
4. Find the mass of the cylinder and the 1.50 mL of water, and record the total mass in Table 1.
5. Now fill the graduate to precisely 3.00 mL and find the total mass once again. Repeat for 5.00 mL.
6. Add an additional 20 drops of water to the cylinder. Estimate the volume to the nearest 0.01 mL, and record it in the data table. (Do not determine the mass of the cylinder and contents.)
7. Finally, add 20 more drops to the cylinder and read the volume once again, estimating to the nearest 0.01 mL.

Cleaning Up

Pour the water down the sink, return the glassware to its proper location, and dispose of the pipet as your teacher directs.



Experiment 2 Report Sheet

Conversion Factors

Name _____

Class _____

Date _____

Prelab Questions

1. Describe the procedure for determining the slope of a straight-line graph.

_____

2. Explain the difference between a *quantity* and a *number.*

_____

_____

3. Every conversion factor can be written in two ways. Complete the table below.

7 days = 1 week	1 day = 0.143 week
12 inches = 1 foot	1 inch = _____ foot
100 cm = 1 m	1 cm = _____ m
$1 = 10 dimes	1 dime = $_____

Data and Observations

Table 1

Mass of empty graduate (step 2):
Mass of graduate + 1.50 mL water (step 4):
Mass of graduate + 3.00 mL water (step 5):
Mass of graduate + 5.00 mL water (step 5):
Volume after adding 20 drops (step 6):
Volume after adding 20 more drops (step 7):

Calculations

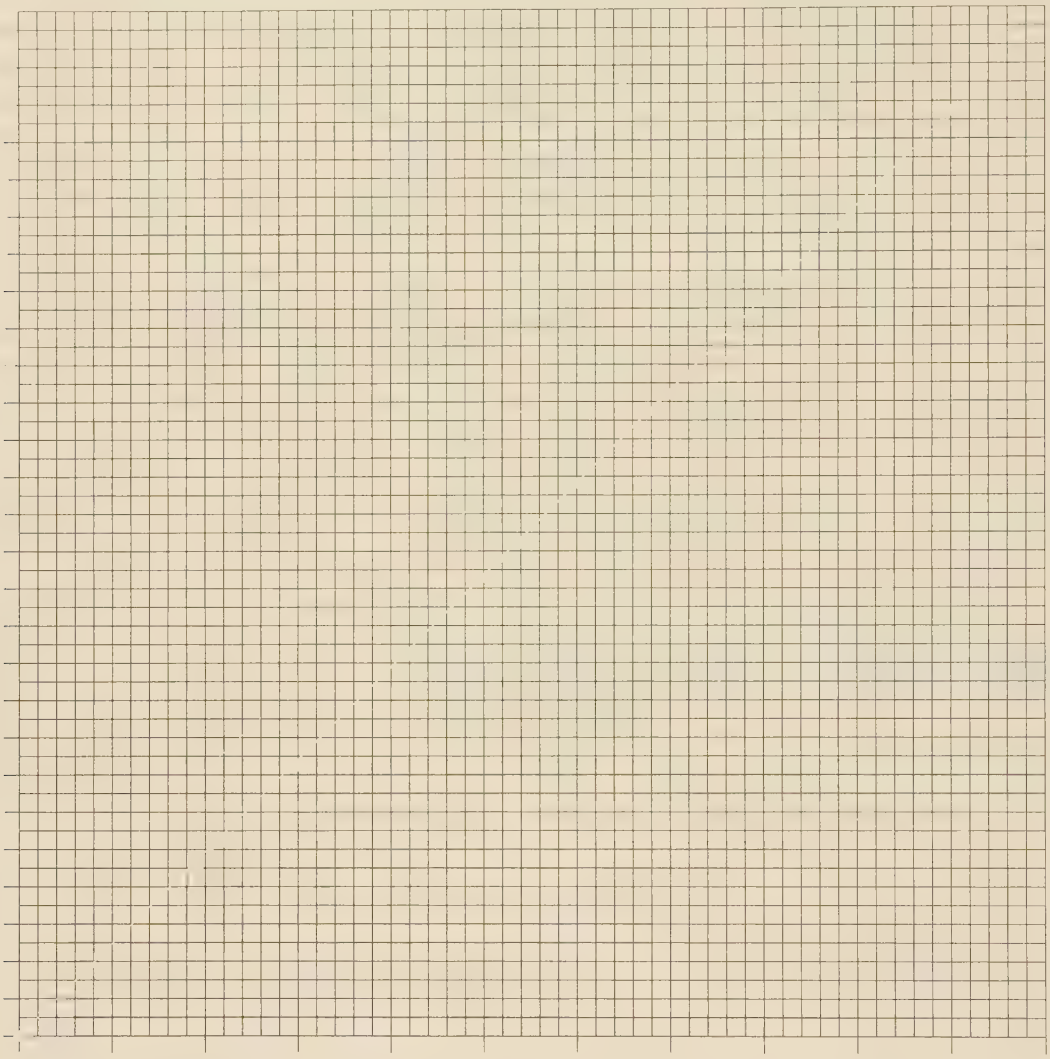
Use the information in Table 1 to calculate the mass of water in 0.00 mL, 1.50 mL, 3.00 mL, and 5.00 mL. Enter the masses in Table 2.

Table 2

Volume of Water	Mass of Water
0.00 mL	
1.50 mL	
3.00 mL	
5.00 mL	

▶ Analysis and Conclusions

1. Use the grid to plot the data from Table 2. Plot volume on the horizontal axis and mass on the vertical axis. Draw the best-fit straight line through the four data points.



2. Graphing your data allows you to find values you didn't actually measure. Use your graph to find (a) the mass of 2.40 mL of water and (b) the volume of 4.25 g of water.

3. Calculate the *slope* of your graph. This gives you the density of water and may be thought of as representing the mass of 1.00 mL of water. Enter the value here and in Table 3.

slope of graph: _____ g/mL

4. Based on your results for steps 6 and 7 of the procedure (see Table 1), (a) what is the average volume of 20 drops of water? What would be the volume of (b) 100 drops? (c) 50 drops? (d) 10 drops? (e) 1 drop? Record your answers to (a) through (e) in Table 3.

5. Use the density you calculated for water to determine (a) the total mass of water in the graduate after the first 20-drop addition and (b) after the second 20-drop addition. What is (c) the average mass of 20 drops of water? Record your answers in Table 3.

6. What is the mass of (a) 1 drop of water? (b) 106 drops ? (c) 6.02×10^{23} drops? Enter the correct masses in Table 3.

Table 3

Density of water (from graph):	_____ g/mL
(4a) Average volume of 20 drops of water:	_____ mL
(4b) Volume of 100 drops of water:	_____ mL
(4c) Volume of 50 drops of water:	_____ mL
(4d) Volume of 10 drops of water:	_____ mL
(4e) Volume of 1 drop of water:	_____ mL
(5a) Mass of water after first 20 drops:	_____ g
(5b) Mass of water after second 20 drops:	_____ g
(5c) Average mass of 20 drops of water:	_____ g
(6a) Mass of 1 drop of water:	_____ g
(6b) Mass of 106 drops of water:	_____ g
(6c) Mass of 6.02×10^{23} drops of water:	_____ g

7. Use your experimental results to complete the following conversions.

a. 1 drop of water = _____ g = _____ mL

b. 1.00 gram of water = _____ mL = _____ drops

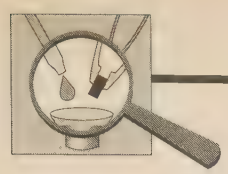
c. 1.00 milliliter of water = _____ g = _____ drops

8. If you did not start with a completely dry graduated cylinder in step 2, how would this affect your answers? Be specific. Will the answers be too high, too low, or unaffected? What part of the answer would be incorrect: the mass? the volume? both?

9. If the drops of water from the pipet hit the sides of the graduate, how would this affect your answers? Again, be specific in describing what could go wrong.

10. The density of water is often assumed to be exactly 1 g/mL. How close to that value did the slope of your graph come? If your value was not in the range of 0.95–1.05 g/mL, was there anything about the procedure that made it difficult for you to get a precise result?

11. Is it reasonable to state "The density of water is exactly 1.00 g/mL."? Explain your answer.



Experiment 3

Classification of Mixtures

You will test four water mixtures to determine whether each is a true solution, a suspension, or a colloid. A basic difference among the three categories is the size of the particles. The particles in a *solution* are too small to be visible, and light will pass through the mixture without being scattered. The particles in a *suspension* are large enough to completely block a beam of light. Initially, you will not be able to see through the mixture. If suspensions are allowed to sit, the large particles settle to the bottom of the container.

Between these two types of mixtures is an intermediate type in which the particles are not large enough to settle or to totally obscure a beam of light, but are large enough to interfere slightly with the passage of a light beam. This interference gives the mixture a cloudy, hazy appearance, which is called the *Tyndall Effect.* (See Figure 2-1 on page 40 of your text.) You are already familiar with this type of mixture in the form of soap suds or haze in the air. Such mixtures are called *colloids.*

You will test four mixtures, then filter each before testing again to see what difference, if any, filtering makes.

OBJECTIVES

1. ***Test*** and ***compare*** mixtures and their filtrates for clarity, settling of suspended particles, and ability to scatter light.
2. ***Classify*** the samples as solutions, suspensions, or colloids.

▶ Materials

Apparatus

2 8- or 12-well test strips
4 15-cm Pasteur pipets
cotton plugs
wooden applicator sticks
safety goggles

Reagents

mixtures A, B, C, D in thin-stem pipets

▶ Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1 and 2 on the Report Sheet.

Procedure

Part 1: Initial Observations

1. Put on your safety goggles.
2. Select one of the four sample pipets, shake it gently, and then place about 8–10 drops into one of the wells in your test strip. The aim is to completely fill the well.
3. Examine the liquid from all sides for evidence of color and clarity. Does the liquid appear homogeneous or heterogeneous? Is it clear or cloudy? If it is cloudy, can you see any evidence of settling? Do not discard the contents of the well; you will want to reexamine it later. Record your observations in Table 1.
4. Repeat steps 2 and 3 with each of the other samples, using a different well for each liquid. In each case, make initial observations, then check again later. The double check is especially helpful in detecting settling.
5. Perform a light scattering (Tyndall Effect) test on each of the four mixtures by placing the test strip over a line of type. A true solution will be clear and will not interfere with your ability to read the letters. With a suspension, the type will be completely obscured. If the mixture is a colloid, it will diminish the sharpness of the letters—the type will be visible, but the letters may seem smaller or blurred. Record the results of your Tyndall Effect tests in Table 1.
6. Based on your observations, classify each sample as a solution, a colloid, or a suspension.

Part 2: Filtration of Mixtures and Testing of Filtrates.

7. You will use a Pasteur pipet and a cotton plug to make a filter as in Figure 1. Prepare four such filters, one for each mixture.

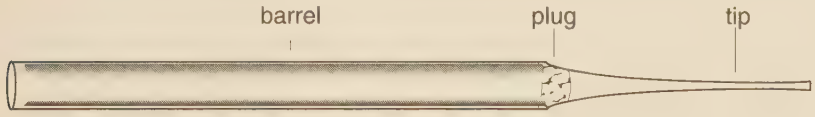
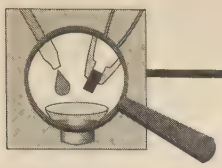


Figure 1: Pasteur pipet with cotton plug

8. Filter the mixtures from the sample pipets through the Pasteur pipets into separate wells of your second test strip. Remember to gently shake each sample pipet before filtering. The liquid that passes through the filter is called the *filtrate.* Try to collect enough filtrate to completely fill a well. After filtration is complete, examine the cotton plug for evidence of particles that were removed by the filtration process. Do not confuse a color on the plug with a solid residue; if you cannot see any solid particles, you were not filtering a suspension.
9. To compare the filtrates with the original samples, repeat steps 3–6. Record any differences you observe in terms of color, clarity, homogeneity, and the ability to transmit or scatter light in Table 2.
10. Based on your observations, classify each filtrate as a solution, a colloid, or a suspension.

▶ ***Cleaning Up*** Return the sample pipets to their proper location. Dispose of the used Pasteur pipets as your teacher directs. Wash the materials in the test strips down the drain; you may need a cotton swab to clean residues from the wells.



Experiment 3 Report Sheet

Classification of Mixtures

Name _____

Class _____

Date _____

Prelab Questions

1. As a general rule, solutions and colloids are homogeneous whereas suspensions are heterogeneous. Define the terms *homogeneous* and *heterogeneous.*

_____

_____

_____

2. Categorize each of these familiar items as a suspension, a colloid, or a true solution. Give a reason for your choice in each case.

a. muddy water

b. lemonade

c. fog

d. milk

e. blood

Data and Observations

Table 1: Initial Observations

Mixture	Appearance	Tyndall Effect (Y/N)	Classification
A			
B			
C			
D			

Table 2: Effect of Filtration

Mixture	Appearance	Tyndall Effect (Y/N)	Particles on Filter (Y/N)	Classification
A				
B				
C				
D				

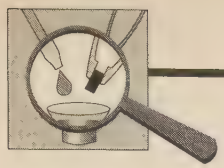
▶ Analysis and Conclusions

1. Did you change the way you classified any of the liquids after filtration? If so, what changes did you make and what did you base them on?

2. What could you do to demonstrate that milk is a colloid even though it doesn't pass the Tyndall Test?

3. Using complete sentences, write a brief summary of your goals for this experiment and of your discoveries. Check your work for spelling and grammar.

4. Compare your observations and conclusions with those of at least three other teams. Then describe and account for any differences. Now that you have completed this activity, you should have a better understanding of the nature of solutions, colloids, and suspensions. Explain how they differ from one another.



Experiment 4

Separation of Mixtures

This investigation will introduce you to three methods for separating mixtures into their components. You will be given a sample of a solid mixture. Your task will be to separate the mixture into three components. You will measure the mass of each component that you recover and compare the total mass recovered to the mass of the original sample.

In principle, one might be able to use a magnifying glass and forceps to painstakingly separate the individual pieces of solid into three separate piles. However, such a procedure is hardly practical, so you will employ more efficient methods. These methods are based on differences in the physical properties of the components.

OBJECTIVES

1. ***Separate*** the components of a solid mixture.
2. ***Determine*** the mass percent of each component in the mixture.
3. ***Compare*** the mass of the recovered components to the mass of the original mixture.

▶ Materials

Apparatus

milligram or centigram balance
10-mL graduated cylinder
10-mL Erlenmeyer flask
boiling stone
filter paper
spatula or wooden stick
hot sand bath
tongs or hot pad (for handling flask)
hand lens (optional)
safety goggles
laboratory apron

Reagents

solid mixture
(in stoppered reaction tube)
distilled water

▶ Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–4 on the Report Sheet.

Procedure

Caution: The sand bath will be very hot (> 200°C) but will not appear so.

1. Put on your safety goggles and laboratory apron.
2. Check that the sand bath is on and the controller is set at 40% power.
3. Find the mass of the stoppered reaction tube containing a sample of the solid mixture and record it in Table 1. When you examine the sample, how many components can you visibly distinguish? Record all observations on the Report Sheet.
4. Add 2 mL of distilled water to the tube. Restopper the tube and shake it vigorously. After a few seconds, feel the tube. Do you notice a temperature change?
5. Find and record the mass of a 10-mL Erlenmeyer flask containing a fresh boiling stone. Carefully transfer all the liquid from the tube into the flask, *without transferring any of the solid.* Add about 1 mL more water to the solid remaining in the tube, restopper, and shake it again. Once again, transfer just the liquid to the Erlenmeyer flask. The process you have just carried out is called *extraction;* the combined liquids that you poured into the flask is called the *extract.* Save this extract for later use.
6. Find and record the mass of a single piece of filter paper. Then place the filter paper on top of a sheet of paper towel that has been folded in fourths. Being careful not to chip or otherwise damage the reaction tube, tap the solid residue from the tube onto the filter paper. Use a spatula or wooden stick to ensure a complete transfer. Label the solid residue and set it aside for later examination.
7. Clean and dry the reaction tube and stopper; determine and record their combined mass. (Note: Even though they look exactly alike, tubes and stoppers are not identical. Be sure that you find the mass of the tube and stopper that you used.)
8. Place your Erlenmeyer flask on the sand bath as in Figure 1. Carefully heat the extract to boiling. The boiling stone prevents boil-over or "bumping" (movement of the reaction vessel from uncontrolled boiling). After the liquid has boiled for about a minute, test the odor of the escaping vapor by gently wafting some of it toward your nose. Maintain a steady rate of boiling while you do step 9.

Caution: Do not directly inhale the vapor.

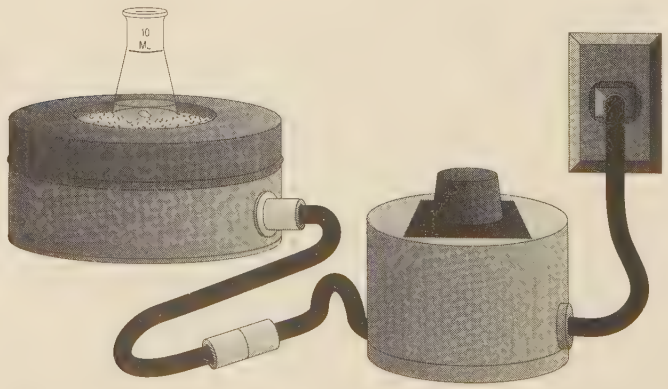


Figure 1

9. Examine the solid on the filter paper and record your observations. Do all the particles appear to be the same substance? How does their appearance compare with particles in the original mixture? Your instructor will provide a safe location where you can put the solid to dry overnight.
10. Continue heating the extract on the hot sand bath until you have evaporated the contents to dryness. No moisture should remain on the walls of the flask. If droplets persist, use tongs to lay the flask on an angle with the lip of the flask on the rim of the sand bath as shown in Figure 2. Turn it occasionally so that all sides are heated, being careful not to get any sand inside the flask. If moisture still persists, ask your teacher for advice.

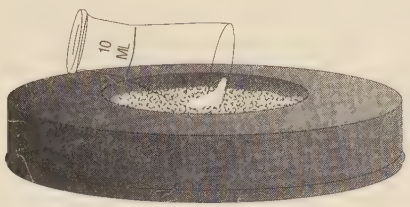


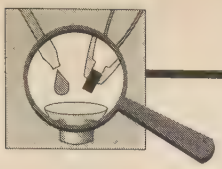
Figure 2

11. When evaporation is complete, use tongs or a hot pad to remove it to a place where it can cool for 5–10 minutes. After the flask has cooled to room temperature, brush off any sand from the outside of the flask. Then measure and record the combined mass of the flask and contents.
12. If one is available, use a hand lens to examine the crystals of solid in the flask. How do they compare to the original mixture? Do they look like the solid that was left after extraction? Record your observations.
13. The following day, measure and record the mass of the filter paper and the insoluble residue.

Caution: Do not put a hot object on the balance. You may damage the balance and your measurements will be inaccurate.

▶ Cleaning Up

The solid left in the Erlenmeyer flask can be safely washed down the drain. Be sure to retrieve the boiling stone, which you can place in the wastebasket. Return the solid left on the filter paper to your teacher for reuse.



Experiment 4 Report Sheet

Separation of Mixtures

Name _____

Class _____

Date _____

Prelab Questions

1. Consult your text or other reference, then write a one-sentence explanation of what is meant by a *volatile* substance.

_____

2. Describe the correct procedure for testing the odors of unknown substances.

_____

_____

_____

3. What physical property will make it possible for you to collect one of the mixture's components by extraction?

_____

4. If you heat a mixture containing one or more volatile components, what is likely to happen?

_____

Data and Observations

Table 1

Mass of stoppered reaction tube and original sample (step 3):

Mass of Erlenmeyer flask and boiling stone (step 5):

Mass of filter paper (step 6):

Mass of empty reaction tube and stopper (step 7):

Mass of Erlenmeyer flask, boiling stone, and soluble residue (step 11):

Mass of filter paper and insoluble residue (step 13):

Observations

Appearance of original mixture (step 3):

Temperature change during dissolving (step 4):

Odor of evaporating solution (step 8):

Appearance of insoluble solid (step 9):

Appearance of residue in reaction tube (step 12):

Calculations

1. Use your data from Table 1 to calculate the first four items in Table 2.

2. Determine the fraction of the original sample recovered (total mass of recovered solids ÷ mass of original sample). Convert this fraction to a percent. Similarly, calculate the percent of the mixture left after extraction and the percent recovered by evaporation. Finally, calculate the percent of material "lost" during the experiment. Record all results in Table 2.

Table 2: Summary of Results

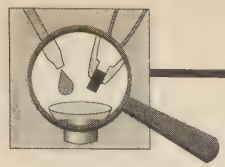
Mass of original sample:
Mass of insoluble residue:
Mass of soluble residue:
Total mass of recovered solids:
Percent of original sample recovered:
Percent of original sample left after extraction:
Percent of original sample recovered by evaporation:
Percent of original sample "lost":

Analysis and Conclusions

1. How does the appearance of each of the recovered solids compare with the original mixture?

2. How many of the three components of the original mixture must have been in the liquid extract? Defend your answer.

3. Account for the material "lost" during the experiment. During what part of the procedure was it most likely "lost"?



Experiment 5

Recognizing Chemical Reactions

Any process in which one or more new substances is formed is called a *chemical change*. Evidence for such a change typically includes one or more of the following: production of a gas (*effervescence*); formation of a solid product (most often seen as a faint cloudiness in what was a clear solution—this solid is called a *precipitate*); an unexpected color change; or a significant increase or decrease in temperature, as heat is either given off (*exothermic reaction*) or absorbed (*endothermic reaction*).

Your task in this experiment is to make accurate observations of any changes that occur when you mix five chemical *reagents* (ree-ay'-junts) in pairs. Your ultimate goal is to decide whether those changes signal the occurrence of a chemical reaction.

OBJECTIVES

1. ***Observe*** and ***record*** the nature of substances, before and after mixing.
2. ***Determine*** whether a reaction takes place between pairs of substances.
3. ***Consult*** with other lab teams, comparing and analyzing your results.

▶ Materials

Apparatus

96-well test plate

tweezers (for handling solid E)

wash bottle

safety goggles

laboratory apron

gloves

Reagents

Solutions A, B, C, D

Solid E

▶ Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–3 on the Report Sheet.

Procedure

1. Put on your safety goggles, laboratory apron, and gloves.
2. At your workstation there is a set of four thin-stem pipets, each containing one of the solutions, A–D. There is also a small container of solid E. Observe these substances and record your observations in Table 1.

Caution: If any of the solutions gets on your skin, wash liberally with water. Because you don't know what the substances are, you must assume they may be harmful.

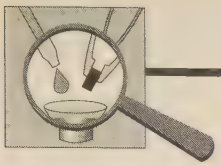
3. Mix three drops of solution A with three drops of solution B in one well of your 96-well test plate. Continue in this fashion until you have tested all possible pairs of solutions, recording your observations in Table 2. (Because you are concerned with possible reactions between pairs of reactants, you should not mix three or more together at one time.)
4. Not all reactions are rapid. If no change is evident at first, observe again after about 5 minutes to see if gradual changes have taken place. To detect any temperature changes, place an ungloved fingertip on the bottom of each well before and after mixing.

Caution: Do not handle any reagent with your bare hands.

5. Next, use tweezers to select individual pieces of E, placing one piece in each of four wells in your test plate. Add 8–10 drops of a different solution to each piece of E. Record your observations in Table 2.
6. When you have completed steps 1–5, compare your results with those of other groups. If you think further tests are needed, complete those tests before cleaning up and returning to the classroom for the postlab discussion.

Cleaning Up

Use your wash bottle to rinse all liquids into the appropriate waste container, as directed by your teacher. Make sure that no solid materials go into the sink. Wash your hands thoroughly before leaving the laboratory.



Experiment 5 Report Sheet

Recognizing Reactions

Name _____

Class _____

Date _____

Prelab Questions

1. What is the distinction between the words *clear* and *colorless* as they apply to solutions? Is a clear solution necessarily colorless? Is a colorless solution necessarily clear?

2. Explain the difference between chemical changes and physical changes.

3. In complete sentences, define the following terms.

   a. precipitate

   b. reagent

   c. exothermic reaction

Data and Observations

Table 1: The Reactants

Substance	Physical State	Description
A	liquid	
B		
C		
D		
E		

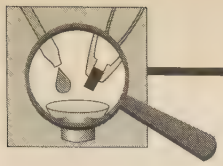
Table 2: The Reactions

Substances	Observations	Reaction (Y/N)	Evidence
A and B			
A and C			
A and D			
A and E			
B and C			
B and D			
B and E			
C and D			
C and E			
D and E			

▶ Analysis and Conclusions

1. How did your results compare with those of other groups? Suggest possible explanations for any significant differences that you observed.

2. If you were to repeat this experiment, how would you change the procedure? Why would you make that change (or those changes)?



Experiment 6

Interpreting Chemical Reactions

When an Alka-Seltzer tablet is dropped into a glass of water, does a chemical reaction take place? How can you tell? There is a lot of fizzing and, although some solid sinks to the bottom of the glass, there is less solid than there was at first. Where did the rest of the solid go? What gas was formed and where did it come from?

When you strike a match, a flame appears and the match seems to disappear. Where did the match go? We could ask the same questions about combustion of propane gas or liquid gasoline. What happens to the material that seems to disappear?

This activity will increase your understanding of chemical reactions. Your task is to make careful observations of the changes that occur when nitric acid reacts with a piece of copper wire. Your teacher will clean the wire with steel wool to remove any tarnish (copper oxide) from the surface. You will relate your observations to the equation for the reaction.

OBJECTIVES

1. ***Observe*** the changes that occur during the reaction of copper and nitric acid.
2. ***Interpret*** the reaction equation in terms of your observations.

▶ Materials

Apparatus

13- × 100-mm culture tube
thin-stem pipet
100-mL beaker
scissors
safety goggles
laboratory apron
gloves

Reagents

3–5 mm copper wire
0.5 mL 7 M nitric acid

▶ Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1 and 2 on the Report Sheet.

Procedure

1. Put on your safety goggles and laboratory apron.
2. Use scissors to cut the bulb of a thin-stem pipet as shown in Figure 1. The top of the bulb and the stem will provide a sealed delivery tube for the safe removal of a toxic gas produced during the reaction.

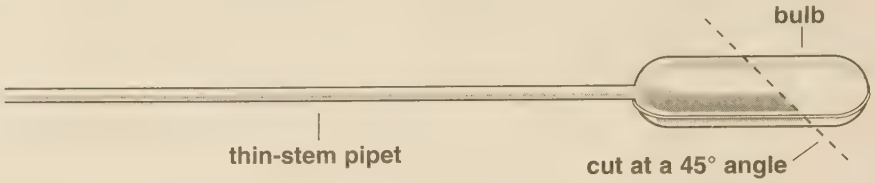


Figure 1: Preparation of a gas delivery tube

Caution: 7 M nitric acid is very corrosive. Avoid all contact with skin and clothing. Wash any spills or splashes off skin and clothing with large amounts of cold running water and notify your teacher immediately.

3. Put on your gloves. Place a 3–5 mm length of copper wire into a 13- × 100-mm culture tube. Fill a 100-mL beaker to the 80-mL mark with tap water. Take the beaker of water, the gas delivery tube, and the culture tube to the fume hood. Your teacher will *carefully* add 7 M nitric acid, HNO_3, to the culture tube to a depth of about 1 cm (about 0.5 mL).
4. *Immediately* fit the bulb of the delivery tube over the mouth of the culture tube and bend the stem into a 180° degree arc as shown in Figure 2. Put the stem end into the beaker of water. You may need to raise the beaker relative to the culture tube to keep the end of the delivery tube under water. When you are certain the system is sealed, *carefully* carry it back to your work station.

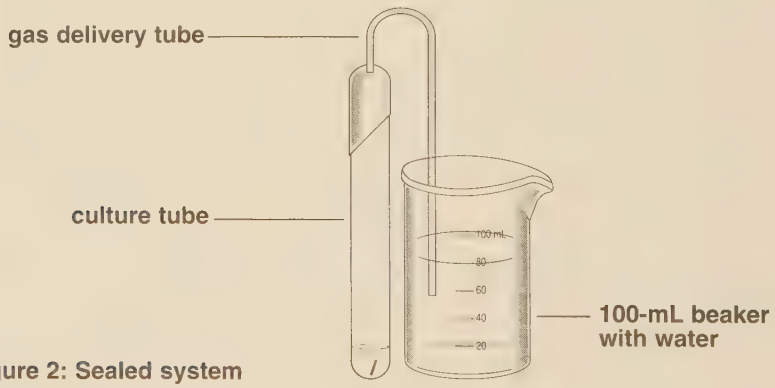
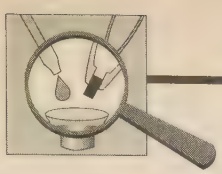


Figure 2: Sealed system

Caution: The gas produced is very toxic. However, it will dissolve in the water before it can escape into the air. The tip of the delivery tube must remain underwater at all times.

5. One partner should hold the culture tube and keep the delivery tube in place; the other partner should hold a piece of white paper behind the culture tube. Record all observations, especially any evidence of reaction, on the Report Sheet. If you have time, begin working on the Analysis and Conclusions questions.
6. When new changes are no longer visible (after about 10 minutes), return your reaction system to the fume hood. You will find a large beaker and a 24-well test plate, both labeled *Experiment 6.* Remove the delivery stem from your culture tube and place it in the beaker. Put the culture tube in the 24-well plate, noting the letter and number that identify the location of your tube. Return to the hood near the end of the period to see if any further changes have occurred.

Cleaning Up Put away the rest of your equipment and wash your hands thoroughly before leaving the laboratory.



Experiment 6 Report Sheet

Interpreting Chemical Reactions

Name _____

Class _____

Date _____

Prelab Questions

1. Use the words *reactants* and *products* in a sentence to show that you understand the meanings of these terms.

_____

_____

2. Predict whether you will see bubbles of gas coming out of the delivery tube during step 5. Offer an explanation for your prediction.

_____

_____

Observations

In the space below, record all the changes you observed during the reaction between nitric acid and copper. Start by describing the appearance of the reactants.

Analysis and Conclusions

There were three products formed by the reaction of copper and nitric acid: an aqueous solution of copper(II) nitrate, nitrogen dioxide gas, and water. The reaction is represented below, first in a word equation, then as a molecular equation. Neither of the equations is balanced.

$$\text{copper} + \text{nitric acid} \longrightarrow \text{copper(II) nitrate} + \text{nitrogen dioxide} + \text{water}$$

$$Cu(s) + HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + NO_2(g) + H_2O(l)$$

For each of the following questions, identify all substances by name and by their symbol or formula, including their physical states—(s), (l), (g), or (aq).

1. Which two products must be in the solution in the culture tube? What color is copper(II) nitrate and how do you know? What became of the copper atoms that were originally part of the wire?

_____

_____

2. What is the name and formula of the orange-brown gas? What elements does it contain? From which reactant did they come?

3. Where did the atoms that formed the water come from?

4. When your teacher cleaned the piece of copper wire before the experiment, what happened to the atoms on the surface of the wire? Was this process a chemical reaction? Explain your answer.

5. In the molecular equation, the same elements are represented on both sides—copper, hydrogen, nitrogen, and oxygen. However, the number of atoms of each element is not the same. To demonstrate this imbalance, complete this atom survey. Three numbers have been entered in the survey to help you get started.

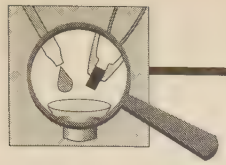
Reactants	**Products**
copper atoms: 1	copper atoms:
hydrogen atoms:	hydrogen atoms:
nitrogen atoms:	nitrogen atoms:
oxygen atoms: 3	oxygen atoms: 9

6. Below is the same molecular equation, but with one major difference: the equation is *balanced.* In other words, the number of atoms of each element is the same on both sides. To show that the equation is now balanced, complete the second atom survey.

$$Cu(s) + 4HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

Reactants	**Products**
copper atoms:	copper atoms:
hydrogen atoms:	hydrogen atoms:
nitrogen atoms:	nitrogen atoms:
oxygen atoms:	oxygen atoms:

7. One way to define a chemical reaction is *a process in which the atoms of the reactants are reorganized.* Explain how this definition applies to the reaction between copper and nitric acid.



Experiment 7

Mass Relationships in Compounds

In this experiment, you will start with an unknown compound. Your task will be to observe and record changes in appearance and mass that accompany a series of reactions. The calculations you will need to do are simple if you collect the proper data. Remember to record all observations and data on the Report Sheet. Because the activity will extend over parts of three class periods, it will be important to label all containers and samples that are stored. Record the date each time you resume work.

OBJECTIVES

1. ***Carry out*** a series of operations on an unknown compound.
2. ***Observe*** changes and ***record*** evidence of chemical reaction.
3. ***Determine*** the changes in mass that accompany the reactions.
4. ***Compare*** data with those of other teams to ***identify*** any regularities.

▶ Materials

Apparatus

milligram or centigram balance
hot sand bath
30-mL or 50-mL beaker
stirring rod
tongs (for handling beaker)
watch glass
forceps or tweezers
filter paper
short-stem funnel
125-mL Erlenmeyer flask
small spatula
safety goggles
laboratory apron
gloves

Reagents

unknown blue compound
aluminum strip, about 5 cm × 0.5 cm
distilled water, in wash bottle

▶ Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–3 on the Report Sheet.

▶ Procedure

Part 1: Day One

1. Put on your safety goggles and laboratory apron.
2. Check that the sand bath is on and the controller is set at 40% power.
3. Find the mass of a labeled, dry 30- or 50-mL beaker. Add between 0.3 and 0.5 g of the unknown blue compound. Record the masses to the nearest 0.001 g in Table 1 on the Report Sheet.

Caution: The sand bath will be very hot (> 200°C) but it will not appear so.

Caution: Do not put a hot object on the balance. You may damage the balance, and your measurements will be inaccurate.

4. Using tongs to hold the beaker steady, gently heat the beaker and sample on a sand bath. Stir constantly with a stirring rod to prevent charring of the solid. (Charring will appear as a black area in the mass of solid.) Continue until the entire sample has changed color. Heating should take less than one minute. Scrape any grains from the stirring rod back into the beaker. Record your observations.
5. Use the tongs to remove the beaker from the sand bath. Allow the beaker and contents to cool for 10 minutes, or until it is at room temperature. Then determine and record the total mass. (To test the temperature of the beaker, place your hand 4–5 cm from the beaker, move it gradually closer, and stop if you detect heat. If you do, wait a minute or two before repeating the test.)
6. Beginning with a few drops at a time, slowly add about 15–20 mL of distilled water to the beaker. Describe all changes in the appearance of the sample as the water is added. Record your observations.
7. Obtain a strip of aluminum metal, shape it into a loose coil, and place it into the solution from step 6. Observe all changes that occur for the first few minutes. Cover the beaker with a watch glass and place it in the location indicated by your instructor until the next laboratory period. Your notes should describe both the changes that occur as the reaction is starting and the appearance of the system the next day.

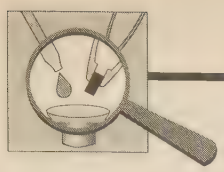
Part 2: Day Two

8. Use tweezers or forceps to remove any aluminum that remains in the beaker. Use a spatula to scrape off any solid clinging to the aluminum into the beaker. Then rinse and discard the aluminum in a container labeled *Aluminum Waste*. Because the solution is acidic, wear gloves and rinse your hands well immediately after you finish.
9. Find the mass of a piece of filter paper, fold it into a cone, and place it in a glass or clear plastic funnel that is suspended in a 125-mL Erlenmeyer flask. Now pour the liquid only from the beaker into the filter, retaining as much as possible of the solid product in the beaker. Allow the liquid to drain through the filter, then rinse the filter lightly with distilled water from a wash bottle.
10. Add about 10–15 mL of distilled water to the solid in the beaker. Use a stirring rod to break up the solid in the water. Drain the liquid through the filter as before, then repeat the rinsing. (In order for all contaminants to be removed, the rinsing must be thorough.)
11. Using a clean spatula, completely transfer your solid product from the beaker to the filter. Use a wash bottle to rinse away any remaining impurities from your product. Drain the rinse water, then repeat. When all the rinse water has drained through the filter, carefully remove the filter with your product from the funnel. Place it on a labeled paper towel in a location where it may be safely left to dry.

▶ ***Cleaning Up*** All solutions may be safely rinsed down the sink with large amounts of water. Wash your hands with soap and water before leaving the laboratory.

Part 3: Day Three

12. Observe and record the appearance of the dry product. Find and record the mass of the dry filter paper and product, then dispose of them as directed by your teacher.



Experiment 7 Report Sheet

Mass Relationships in Compounds

Name _____

Class _____

Date _____

Prelab Questions

1. Identify the type of reaction discussed in Chapter 3 that would best describe the following situation: A reaction begins with a single reactant and results in a apparent loss of mass.

_____

_____

2. Why must you avoid mixing the solid product with water while it is in the filter?

_____

_____

3. Discuss the effects that might result from the following examples of typical student errors in technique. In each case, specify whether the mass of the sample will appear to be greater or less than it should be.

   a. A student forgets to find the mass of the empty beaker before beginning, and finds the mass of another beaker.

_____

_____

_____

   b. A student finds the mass of the empty beaker, then adds a masking-tape label for ease of identification.

_____

_____

_____

   c. A student places the stirring rod used to mix the solid and water on the desk top between uses.

_____

_____

_____

   d. To save time, a student finds the mass of the filter and product immediately after filtration, rather than waiting for it to dry overnight.

_____

_____

_____

Data and Observations

Table 1

Mass of labeled beaker:

Mass of beaker and unknown sample:

Mass of beaker and sample after heating:

Mass of filter paper:

Mass of filter and dried product:

Observations

Appearance of original sample:

Observations during heating (step 4):

Observations during addition of water (step 6):

Observations during reaction with aluminum (step 7):

Appearance of final product (step 12):

Calculations

Show your work and enter the results in Table 2.

1. Calculate the mass of solid remaining in the beaker after heating (step 5).

2. Calculate the mass of the final product (step 12).

3. Calculate the ratio of the mass of the solid remaining after heating to the mass of the original blue solid.

4. Calculate the ratio of the mass of final product to the mass of the original blue solid.

Table 2

Mass of original sample:

Mass of solid after heating:

Mass of final product:

Mass of solid after heating ÷ mass of original sample:

Mass of final product ÷ mass of original sample:

▶ Analysis and Conclusions

1. Given what you observed when you added distilled water in step 6, suggest a reasonable explanation for the loss of mass during heating.

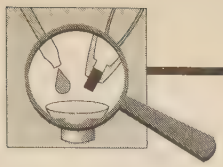
2. What might have happened if you had left the beaker and contents from step 5 overnight before finding the mass?

3. Cite at least three examples of evidence for chemical change that you observed in step 7.

4. Of the categories of reactions discussed in Chapter 3, which type was involved in step 7? Explain your answer.

5. Does your final product appear to be the same substance as either of the other solids in this experiment? Explain.

6. Compare your ratios from calculations 3 and 4 to those of other teams. Are the results consistent or do they appear to vary in random fashion? Does this experiment illustrate the Law of Definite Composition? Explain.



Experiment 8

An Investigation of the Mole

We often find it convenient to deal with certain types of objects in groups, rather than individually: eggs and doughnuts are bought and sold in groups of twelve, which we call dozens; days are grouped into weeks, each consisting of seven days; tennis balls and golf balls come in containers of three.

In 1811, Amadeo Avogadro hypothesized that equal volumes of gases at the same temperature and pressure contain the same number of particles of each gas. Thus, if an oxygen sample is 16 times as massive as an equal volume of hydrogen, then necessarily *each molecule* of oxygen is 16 times as massive as *each molecule* of hydrogen. Through such comparisons, chemists were able to establish a set of *relative atomic masses* for the elements.

These atomic masses had no practical use because no one could isolate an individual atom. Chemists needed to group atoms into a practical standard; they chose a unit called the *mole.* (The word *mole* is an abbreviation of the German word *molekular.*) One mole of an element is the number of atoms of that element that has a mass in grams numerically equal to the element's relative atomic mass. One mole is equal to 6.02×10^{23} items.

OBJECTIVES

1. ***Establish*** a standard quantity of beans that is analogous to a *mole* of atoms.
2. ***Simulate*** the process for establishing relative atomic masses.

▶ Materials

Apparatus

centigram or milligram balance

small plastic cups

safety goggles

Reagents

dried beans of at least 5 types

▶ Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1 and 2 on the Report Sheet.

▶ Procedure

Part 1: Determine Mass for 50-bean Samples

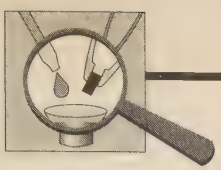
1. Put on your safety goggles.
2. Find the mass of an empty cup (±0.01 g); that is, report its mass to the nearest one-hundredth of a gram. Enter the mass in Table 1 on your Report Sheet.
3. Fill the cup with exactly 50 beans of one type. Discard any beans that are not "typical." Find and record the mass of the cup with the 50 beans. Then subtract to find the mass of the 50 beans.
4. Repeat step 3 for each remaining type of bean.
5. Complete the Calculations section on the Report Sheet and then do Part 2 of the procedure.

Part 2: Number of Beans Equal to Relative Mass in Grams

6. For each type of bean, measure out a mass in grams equal to the relative mass that you calculated for that bean. In each case, determine the whole number of beans that comes closest to the relative mass. (The actual mass may be slightly below or slightly above the mass you want.) Record your data in Table 2 in the column labeled *Number in 1 Pot.* We have arbitrarily defined a *pot* of beans as the number of beans that has a mass in grams equal to the relative mass of that type of bean.

▶ Cleaning Up

Return all beans to the large containers, being careful not to contaminate one type of bean with beans of another type. Do not discard the measuring cups.



Experiment 8 Report Sheet

An Investigation of the Mole

Name _____

Class _____

Date _____

Prelab Questions

1. What number is represented by each of the following terms? Consult a dictionary for any that you do not recognize.

   a. score

   b. gross

   c. decade

   d. century

2. How many objects are indicated in each of the following quantities?

   a. 7 dozen eggs

   b. 0.25 gross of pencils

Data and Observations

Table 1

Bean Type	Symbol	Mass of Cup and 50 Beans	Mass of 50 Beans
		Mass of empty cup: _____ g	

Calculations

1. Calculate the relative mass of each type of bean. Find the relative mass by dividing the sample mass (Mass of 50 Beans) by the sample mass of the least massive bean. The least massive bean will have a relative mass of 1.00; all other beans will have a relative mass greater than 1.00. (It is not necessary to show all calculations, but you should show at least one sample calculation of relative mass.) Record your results in Table 2.

2. Complete Table 2 by ranking the beans from least massive to most massive, starting with a rank of 1 for the bean with a relative mass of 1.00.

Table 2

Bean Type	Symbol	Relative Mass	Rank	Number in 1 Pot

Analysis and Conclusions

1. Using the data in Table 2, compare the numbers of beans in a pot for the various types. Allowing for the fact that only whole numbers of beans could be used and for experimental errors, were the results fairly consistent from type to type? What variation, if any, did you find?

2. Calculate the average number of beans in a pot and express your answer with an uncertainty that reflects the range of variation. Suppose you were averaging the numbers 26, 28, 29, 29, 28. You would report the average as 28 ± 2 to indicate that none of the numbers being averaged is more than 2 units above or below the average.

3. Calculate the following quantities for each type of bean, showing the setups for one example of each of the five calculations, a–e.

   a. number of pots in 250 grams

   b. number of beans in 250 grams

   c. number of pots in 250 beans

   d. number of beans in 3.17 pots

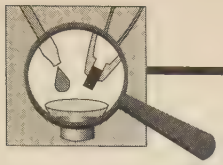
   e. number of grams in 3.17 pots

4. For some quantities, the result is the same no matter which bean is used; for other quantities, different beans give different results. What accounts for this difference?

5. Why should answers to 3b and 3d be reported as whole numbers? How does this restriction against using a fractional number of beans parallel our understanding of atoms?

6. In a single paragraph, explain how this experiment simulates the calculation of relative atomic masses shown on the periodic table and how the pot provides an analogy to the mole.

7. The mass of a box is determined before and after the box is filled with 100 glass marbles. The marbles are replaced by 100 plastic spheres and the mass is determined again. Explain how you would calculate the mass ratio of one glass marble to one plastic sphere.



Experiment 9

Identifying an Unknown Oxide

The goal of this investigation is to determine whether a copper oxide sample is copper(I) oxide, Cu_2O, or copper(II) oxide, CuO. To begin, you will dissolve the sample in hydrochloric acid, converting copper oxide to copper chloride. The second product is water.

Once the oxide has completely reacted, you will add a small piece of aluminum, which will react with the solution, producing copper metal and an aqueous aluminum chloride solution, $AlCl_3(aq)$.

You will dry the copper and determine its mass. From the mass of copper and the mass of the copper oxide sample, you will be able to calculate the percent of copper in the original sample. Because the accepted values for the percent of copper in copper(I) oxide and in copper(II) oxide are significantly different, it should be evident which oxide you started with.

OBJECTIVES

1. ***Determine*** the percent of copper in an unknown copper oxide.
2. ***Compare*** the percent composition of the unknown with the percent compositions of copper(I) oxide and copper(II) oxide.

▶ Materials

Apparatus

milligram or centigram balance
30–100 mL beaker
watch glass
steel wool
paper towel
forceps or tweezers
spatula
stirring rod
filter paper
short-stem funnel
125-mL Erlenmeyer flask
laboratory apron
gloves
safety goggles

Reagents

0.5 g of unknown copper oxide
5 mL of 6 M HCl
aluminum strip,
about 0.5 cm × 5 cm
distilled water in wash bottle

▶ Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1 and 2.

Caution: This experiment generates fumes that may cause discomfort to contact lens wearers.

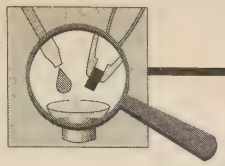
Caution: 6 M HCl is corrosive. Wash spills on skin with large quantities of cold running tap water. Remove contaminated clothing immediately. All spills should be reported to the teacher. Keep your face away from the beaker while the acid and oxide are reacting. If the room is not well ventilated or solution starts to fizz out of the beaker, cover the beaker loosely with a watch glass and set it in the hood.

▶ If you cannot complete the lab in one period, stop after step 6. Cover the beaker with plastic wrap and place it in a well-ventilated, secure location.

▶ ***Cleaning Up*** The filtrate may be safely washed down the drain. Follow your teacher's instructions for disposal of the solid copper product. The filter paper may be placed in the wastebasket once the copper has been removed.

▶ Procedure

1. Put on your laboratory apron, gloves, and safety goggles.
2. Label and find the mass (±0.001 g) of a clean, dry small beaker. Add 0.4–0.6 g of the unknown oxide from the supply bottle. Then determine the mass of the beaker and sample (±0.001 g). Record all measurements in the Data Table on the Report Sheet.
3. Carefully add an estimated 5 mL of 6 M HCl to the beaker, then gently swirl the beaker to ensure that the acid makes good contact with the copper oxide. Hold a piece of white paper behind the beaker while you observe any color changes.
4. While waiting, polish a strip of aluminum with fine steel wool to remove the thin coating of aluminum oxide from the surface. Wipe the metal clean with a paper towel, then form it into a loose coil.
5. When the reaction appears to be complete (all of the oxide has reacted), slowly and carefully add distilled water, 1 or 2 mL at a time, until you have added a total of about 10 mL. Swirl the beaker after each addition. Then place the coiled aluminum strip in the solution. If the liquid is not deep enough to cover the metal, add just enough water so the coil is covered, and observe the reaction that occurs.
6. After 5 minutes, use forceps to hold the strip and a spatula to scrape the solid that has formed into the beaker. Allow the strip to react for another 2–3 minutes and repeat the scraping. When no new solid forms on the cleaned portion of the coil, the reaction is complete.
7. When the reaction is complete, use forceps to remove the unreacted aluminum, carefully scraping any remaining copper back into the beaker. Then decant (pour off) as much liquid as you can without loss of solid product.
8. Add about 10 mL of distilled water to your product, swirl the beaker and contents. Allow the solid to settle and decant the liquid once again. Add a second 10-mL portion of distilled water to the solid, but this time use a stirring rod or spatula to thoroughly mix the water with the solid; this will remove any aluminum chloride that may remain. As before, allow the solid to settle before decanting.
9. While waiting for the solid to settle, label a piece of filter paper and then find and record its mass. Fold the filter paper and place it in a funnel that is resting in a 125-mL Erlenmeyer flask. You will use this setup to collect your copper product by filtration.
10. Pour just the water through the filter first, then use your stirring rod or spatula to completely transfer all of the solid to the filter. You may need to use water from your wash bottle to help in the transfer. Rinse the filter with three successive portions of distilled water, then allow all the liquid to drain through as fully as possible.
11. Carefully remove the filter paper and product from the funnel to a place designated by your teacher. Place it on a folded paper towel to dry until the next lab period, at which time you will determine the mass of the filter and solid.



Experiment 9 Report Sheet

Identifying an Unknown Oxide

Name _____

Class _____

Date _____

Prelab Questions

1. Write balanced equations for the reactions described in the introduction.

a. copper(I) oxide and HCl

b. copper(II) oxide and HCl

c. copper(I) chloride and Al

d. copper(II) chloride and Al

2. Calculate the percent of copper in copper(I) oxide and in copper(II) oxide. Show your calculations in the space provided.

Data and Observations

Data Table

Mass of empty beaker (step 2):

Mass of beaker and copper oxide (step 2):

Mass of labeled filter paper (step 9):

Mass of filter paper and copper product (step 11):

▶ Calculations

1. Calculate the mass of copper oxide used and the mass of copper metal recovered. Show your calculations in the space below.

mass of copper oxide: _____ g

mass of copper metal: _____ g

2. Calculate the percent of copper in your copper oxide sample, then refer back to Prelab question 2 to decide whether the unknown is copper(I) oxide or copper(II) oxide.

Percent of Copper: _____ %

Identity of Unknown: _____

3. Calculate your percent error, according to this formula:

(experimental % Cu) – (accepted % Cu) ÷ (accepted % Cu) × 100 = percent error

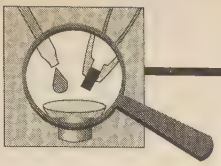
A negative percent will mean that your result was lower than the accepted value.

Error = _____ %

4. Calculate the mass of copper that should have been recovered, using the mass of your copper oxide sample and the accepted value for the percent of copper in that oxide. Show your calculations. (Hint: Convert percent to a decimal for the calculations.)

▶ Analysis

Prepare for the postlab discussion by reviewing the procedure. List the factors that might have contributed to an error in the amount of copper you recovered.



Experiment 10

Synthesis of Manganese Chloride

In this experiment you will synthesize manganese(II) chloride by reacting a known mass of manganese metal with an excess of hydrochloric acid. To calculate the percent yield, you will compare the theoretical yield to the actual mass of product recovered. The balanced equation for the reaction is:

$$\underset{54.9\ g}{Mn(s)} + 2\ HCl(aq) \longrightarrow \underset{125.9\ g}{MnCl_2(aq)} + H_2(g)$$

Theoretically 54.9 grams of manganese will generate precisely 125.9 grams of manganese(II) chloride. These relative masses must apply regardless of the mass of manganese used.

OBJECTIVES

1. ***Synthesize*** manganese(II) chloride using manganese metal and hydrochloric acid.
2. ***Determine*** theoretical yield based on the mass of manganese used.
3. ***Determine*** the percent yield by comparing the theoretical yield with the actual yield.

▶ Materials

Apparatus

milligram or centigram balance
10- or 25-mL Erlenmeyer flask
hot sand bath
tongs or hot pad (for handling flask)
desiccator
safety goggles
laboratory apron
gloves

Reagents

0.1 g manganese metal
1.5 mL 6 M hydrochloric acid
boiling stone

▶ Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–5 on the Report Sheet.

▶ Procedure

1. Put on your laboratory apron, safety goggles, and gloves.
2. Check that the sand bath is on and the controller set at 40% power.
3. Add a small boiling stone to a labeled 10- or 25-mL Erlenmeyer flask and determine the mass of the flask. Place about 0.05–0.1 g of manganese metal in the flask; find the mass of the flask and metal. Record both masses in your Data Table.

Caution: The sand bath will be very hot (> 200°C) but will not appear so.

Caution: 6 M HCl is corrosive to skin and clothing. Use only in the fume hood. Because hydrogen is flammable, avoid use of any burners during the reaction.

Caution: HCl vapor is given off during evaporation. If the room is not well ventilated, evaporation should be done in the hood.

4. Bring your flask to the fume hood and add about 1.5 mL of 6 M hydrochloric acid, HCl(aq).
5. Allow the acid to react with the metal until the sample has reacted completely (about 5 minutes). All the manganese must be converted to manganese(II) chloride before you begin the evaporation. During the reaction, try to determine whether the reaction is exothermic or endothermic.
6. Place the flask on the sand bath to evaporate the solvent as shown in Figure 1a. Continue until the water has entirely evaporated, leaving only a pinkish-white solid residue. There should be no moisture on the walls of the flask. If droplets persist, use tongs to lay the flask carefully on an angle with the lip of the flask on the rim of the sand bath as shown in Figure 1b.

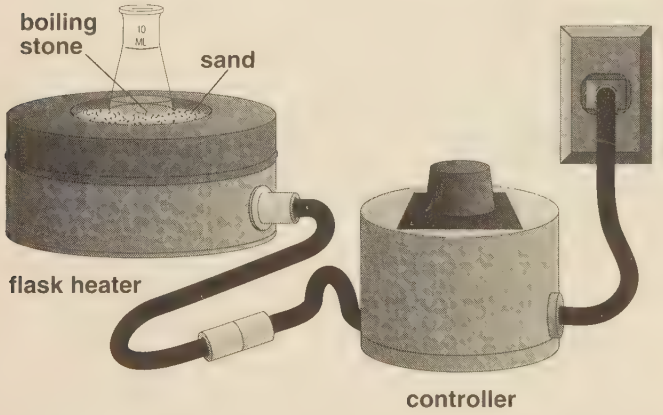


Figure 1a

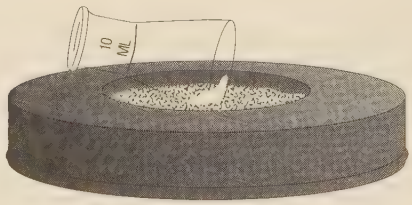
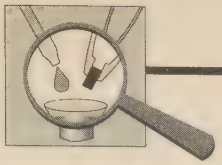


Figure 1b

7. When the flask is thoroughly dry, use tongs or a hot pad to remove it from the sand bath to a place where it can cool for 5–10 minutes. A desiccator made from a large coffee can with a drying agent in the bottom is ideal. The drying agent keeps the product from absorbing moisture from the air.
8. After the flask has cooled to room temperature, determine the mass of the flask and product.
9. Return the flask to the hot sand bath for an additional 3–5 minutes heating. Cool the flask in the desiccator, then measure its mass and record in your Data Table. Ideally, the two measurements should agree to within 0.003 g.
10. If the mass has changed by more than 0.005 g, repeat step 9. Continue until you get two successive measurements that agree within 0.005 g.

▶ Cleaning Up

Hand your labeled flask in to your teacher. Manganese compounds require special handling for disposal; do not wash them down the sink. Before leaving the lab, clean up all materials and wash your hands thoroughly.



Experiment 10 Report Sheet

Synthesis of Manganese Chloride

Name _____

Class _____

Date _____

Prelab Questions

1. Explain the difference between *theoretical* yields and *actual* yields.

2. Why is it important for the manganese to react entirely before proceeding?

3. What is a desiccator, and why is it needed in this experiment?

4. If 6 M HCl accidentally spills or splashes on your skin or clothing, describe in detail what you should do. What, if anything, would you do differently if acid is spilled on the surface of the lab bench or hood?

5. When you are done with the manganese(II) chloride, how should you dispose of this product?

Data and Observations

Data Table

Mass of empty Erlenmeyer flask (with boiling stone):

Mass of Erlenmeyer flask and manganese metal:

Mass of flask and manganese(II) chloride (step 8):

Mass of flask and manganese(II) chloride (step 9):

Mass of flask and manganese(II) chloride (step 10):

▶ Calculations

If necessary, use a separate piece of paper for calculations.

1. Calculate the mass and the number of moles of manganese metal used.

Mass of Mn used _____ g

Moles of Mn used _____ mol

2. Use the final mass of the flask and product to find the mass and number of moles of manganese(II) chloride produced. This is your actual yield.

Mass of $MnCl_2$ _____ g

Moles of $MnCl_2$ _____ mol

3. Calculate the mass of manganese(II) chloride you would expect to produce from the amount of manganese used.

Theoretical yield of $MnCl_2$ _____ g

4. Determine the percent yield of manganese(II) chloride.

$$\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100 = \text{Percent Yield}$$

Percent yield of $MnCl_2$ _____ %

▶ Analysis and Conclusions

1. Predict the effect on your yield for the following procedural errors. In each case, will the reported mass of manganese(II) chloride be too high or too low? Give a reason for each answer.

a. The manganese metal used was impure; it had a coating of manganese oxide.

_____

b. Some of the solution was lost during evaporation because of spattering.

_____

c. The water was not completely evaporated.

_____

d. During cooling, the flask was left in the open air instead of in a desiccator.

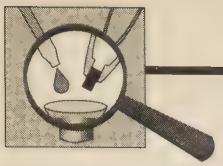
_____

2. Was your percent yield greater than or less than 100%? Suggest one or more possible explanations for any variation from 100% yield. You are not limited to the choices given in question 1. Base your answer on your own experience doing the experiment.

_____

_____

_____



Experiment 11

Gas Laws I

Experimental measurements on a variety of gases show that, under the same conditions of temperature and pressure, equal quantities of all gases have the same volume—knowledge you will use in Experiments 11 and 12. In this experiment, you will react magnesium metal with an excess of hydrochloric acid, HCl, to produce hydrogen gas.

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

You will collect the hydrogen by the displacement of water in a 10-mL graduated cylinder. Because you will be using a very small amount of magnesium, you will need to be extremely careful to justify a precision of two significant figures. Your instructor will supply the mass of a 100.–cm sample of magnesium ribbon and the current barometric pressure, which you will need for your calculations.

OBJECTIVES

1. ***Measure*** the volume of hydrogen gas produced by complete reaction of a known length of magnesium ribbon with an excess quantity of hydrochloric acid.
2. ***Determine*** the volume of a mole of hydrogen at laboratory conditions of temperature and pressure.
3. ***Calculate*** the volume of a mole of hydrogen at standard temperature and pressure (STP).

▶ Materials

Apparatus

400-mL beaker
10-mL graduated cylinder
#00 1-hole rubber stopper
copper wire
metric ruler
wash bottle with distilled water
thermometer, –20° to + 110°C
plastic teaspoon
safety goggles
laboratory apron
gloves

Reagents

magnesium ribbon
3 M HCl
baking soda, $NaHCO_3$

Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–3 on the Report Sheet.

Procedure

1. Put on safety goggles, a laboratory apron, and gloves.
2. Fill a 400-mL beaker about three-fourths full of tap water; the water should be at or near room temperature.
3. Obtain a small piece of magnesium ribbon (about 0.8–1 cm in length). Measure and record its length to the nearest 0.01 cm. Use a 10-cm piece of copper wire to make a "cage" for your magnesium. Make a hook at one end of the copper wire and slide the magnesium inside. Then fold each "tab" of magnesium over one leg of the hook as shown in Figure 1a. Press the cage as tight as possible and wrap the copper wire once around the cage in one tight 360° loop as shown in Figure 1b.

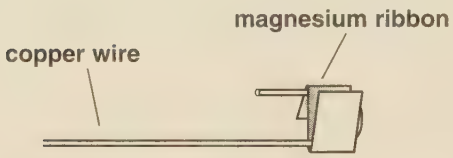


Figure 1a

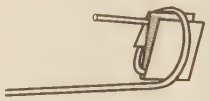


Figure 1b

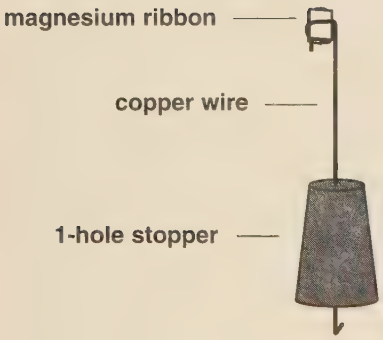


Figure 2: The stopper assembly

4. Fit the wire cage into a #00 1-hole rubber stopper as shown in Figure 2. The cage should be about 2–3 cm from the small end of the stopper. When you have the cage at the desired distance from the stopper, make a right-angle bend in the wire where it exits the top of the stopper.
5. Carefully add about 3 mL of 3 M HCl from a dropper bottle to a 10-mL graduated cylinder. Using a wash bottle, slowly add distilled water to the graduated cylinder until it is completely full. Direct the water down the side of the graduate to minimize mixing of the water and the acid. Insert the stopper assembly into the top of the cylinder. Water should escape through the hole in the stopper; if it does not, remove the stopper and add more water before replacing the stopper. This procedure ensures that no air is trapped in the cylinder.

   Note: Adding water to acid in this procedure is an exception to the general rule that acid must be added to water to avoid splashing of the acid.

Caution: 3 M hydrochloric acid is corrosive. Avoid all contact with skin and clothing. If any gets into your eyes, use the eyewash fountain immediately. Wash any spills off skin or clothing with plenty of water and notify your teacher.

6. Place your finger over the hole in the stopper and invert the graduated cylinder, lowering it into the beaker of water. Remove your finger when the stopper is below the level of water in the beaker. The hydrochloric acid is denser than pure water, so it will slowly sink toward the stopper and the magnesium. Observe and record evidence of reaction. (See Figure 3.)

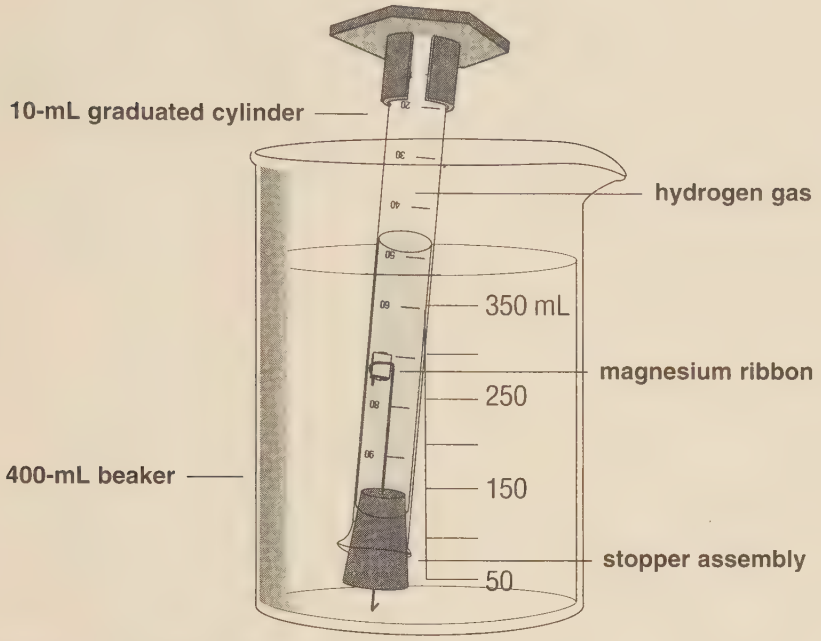


Figure 3: The reaction setup

7. When reaction is complete, allow the system to stand for two or three minutes. Tap the sides of the graduate to dislodge any gas bubbles that may be clinging to the glass wall. Make sure that there are no small pieces of unreacted magnesium on the wall of the cylinder. If a small piece remains, gently shake the graduate up and down to wash the metal back into the acid solution, allowing it to finish reacting. (Be careful not to lift the graduate completely out of the water.)

8. Move the graduate up or down until the water levels inside and out are even to ensure that the total pressure on the gases inside the cylinder is the same as the barometric pressure in the room. Record the volume of gas trapped in the cylinder (±0.1 mL). (Remember that you are reading marks on an inverted cylinder.) Measure the temperature of the water in the beaker. The temperature of the water is assumed to be the same as the temperature of the trapped gases. Enter these data in Table 1 on the Report Sheet.

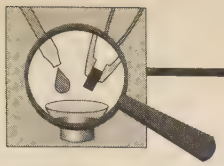
9. Clean up and then repeat steps 2–8 for two additional trials.

▶ Cleaning Up

Excess hydrochloric acid should not be poured down the drain. The acidic solution in the beaker can be neutralized with sodium hydrogen carbonate (ordinary baking soda) to produce a carbonated solution of table salt.

$$HCl(aq) + NaHCO_3(s) \longrightarrow H_2O(l) + CO_2(g) + NaCl(aq)$$

As part of the Prelab, you calculated the mass of $NaHCO_3$ needed to neutralize 3 mL of 3 M HCl. Use a plastic teaspoon to add approximately that amount of the solid (a little at a time to minimize foaming) to the beaker. Stir well and flush the solution down the drain.



Experiment 11 Report Sheet

Gas Laws I

Name _____

Class _____

Date _____

Prelab Questions

1. Explain how the mass of a piece of magnesium ribbon may be calculated from its length and the mass of a 100.-cm strip of the ribbon.

_____

2. Use Dalton's Law of Partial Pressures to explain why the pressure of hydrogen gas will be less than the observed barometric pressure in the laboratory. How will you determine the pressure of the hydrogen you produce?

_____

_____

_____

_____

3. Calculate the mass of baking soda, $NaHCO_3$, needed to neutralize 3.0 mL of 3.0 M hydrochloric acid, HCl. Show your calculations. (Hint: The amount needed is less than 1 gram.)

_____ g $NaHCO_3$

Data and Observations

Table 1

mass of 100. cm of magnesium ribbon _____ g

barometric pressure _____ kPa

	Trial 1	Trial 2	Trial 3
length of magnesium ribbon	_____ cm	_____ cm	_____ cm
volume of hydrogen collected	_____ mL	_____ mL	_____ mL
temperature of water	_____ °C	_____ °C	_____ °C
vapor pressure of water at measured temperature	_____ kPa	_____ kPa	_____ kPa

▶ Calculations

1. Calculate the mass of magnesium used in each trial, using the mass of 100. cm of ribbon and record your answers in Table 2.

2. Calculate (a) the number of moles of magnesium used for each trial and (b) the number of moles of hydrogen gas produced. Enter the values in Table 2.

3. Because you collected hydrogen over water, some of the gas in the cylinder at the end of the reaction is water vapor; we say the hydrogen gas is "wet." The amount of water that evaporates is dependent only on the temperature. Use Table 3 to find the pressure due to water vapor in the graduate. Then subtract this value from the barometric pressure to get the pressure due to "dry" hydrogen. Be sure to report the pressure to the correct degree of precision.

4. You measured the volume occupied by a very small fraction of a mole of hydrogen at specific pressure and temperature conditions. For each trial, calculate the volume that 1.00 mole of hydrogen would occupy at your experimental temperature and pressure (called *laboratory conditions*).

5. Use Charles's Law to calculate the volume that 1.00 mole of hydrogen would occup at your experimental pressure, but at 0°C, which is called *standard temperature.*

6. Use Boyle's Law to calculate the volume that one mole of hydrogen would occupy at your experimental temperature, but at 101.3 kPa, which is called *standard pressure.*

7. Calculate the volume that 1.00 mole of hydrogen would occupy at 273 K and 101.3 kPa, which are standard temperature and pressure (STP). This volume is called the *standard molar volume,* which you will learn more about in Chapter 7.

Table 2

	Trial 1	Trial 2	Trial 3
mass of magnesium ribbon used	_____ g	_____ g	_____ g
moles of magnesium used	_____ mol	_____ mol	_____ mol
moles of hydrogen gas produced	_____ mol	_____ mol	_____ mol
pressure of dry hydrogen	_____ kPa	_____ kPa	_____ kPa
molar volume of H_2 (lab conditions)	_____ L/mol	_____ L/mol	_____ L/mol
volume of 1 mol H_2 (at 0°C)	_____ L	_____ L	_____ L
volume of 1 mol H_2 (at 101.3 kPa)	_____ L	_____ L	_____ L
molar volume of H_2 (STP)	_____ L	_____ L	_____ L

Table 3: Vapor Pressure of Water

Temperature (K)	Vapor Pressure (kPa)	Temperature (K)	Vapor Pressure (kPa)
283	1.23	294	2.49
284	1.31	295	2.64
285	1.40	296	2.81
286	1.50	297	2.98
287	1.60	298	3.17
288	1.71	299	3.36
289	1.82	300	3.56
290	1.94	301	3.78
291	2.06	302	4.01
292	2.20	303	4.24
293	2.34		

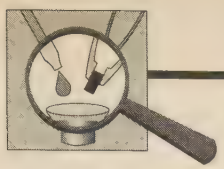
Not all barometers are calibrated in kilopascals, kPa. These conversion factors may be helpful:
101.33 kPa = 760 torr =
760 mm Hg = 1.00 atm

▶ Analysis and Conclusions

1. What property of hydrogen gas makes it possible for you to collect it in your graduated cylinder?

2. Predict the molar volume of oxygen collected at STP. Justify your answer.

3. You used a 10-mL graduated cylinder and less than 1.00 cm of magnesium ribbon. This experiment is often done using 5.0 cm of magnesium ribbon and a 50-mL tube, graduated in tenths of a milliliter. Suggest two reasons why the larger-scale experiment might produce more meaningful results.



Gas Laws II: The Universal Gas Constant

In Experiment 11 you measured the volume of a gas at laboratory conditions and used it to determine the volume that a mole of gas would occupy at STP. To complete the calculations, you used your knowledge of Boyle's and Charles's Laws. The variables you used can be combined into a single equation

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

where P_1, V_1, and T_1 refer to an initial set of conditions and P_2, V_2, and T_2 refer to the final conditions. What the equation tells us is that the value of the expression, *PV/T,* does not change. If that is the case, then we can determine a value for *PV/T* that is valid under any set of conditions—as long as the sample remains a gas and as long as the quantity of gas does not change. It would be even more useful to have a general value that is valid for any quantity of gas.

If we introduce a new variable, *n,* for the number of moles of gas in the system, we can expand the expression to

$$\frac{PV}{nT} = \text{a constant}$$

This relationship is known as the *ideal gas law,* which is usually written as $PV = nRT$. The constant, *R,* is called the *universal gas constant.*

You will use your data from Experiment 11 to determine a value of *R*. The units for *R* must reflect pressure times volume, divided by number of moles times temperature. You will be using SI units, so *R* will be in kPa•L/mol•K. You will calculate the value of *R* for each of three trials, as well as an average result which will then be used to determine your percent error.

OBJECTIVES

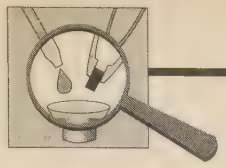
1. ***Determine*** a value for the universal gas constant, *R*.
2. ***Compare*** your experimental value for *R* to the accepted value for *R* and ***calculate*** a percent error.

Prelab

1. Read the introduction and procedure before your begin.
2. Answer Prelab questions 1–4 on the Report Sheet.

Procedure

Using your data from Experiment 11, enter the following values for hydrogen into the Summary Table on the Report Sheet: number of moles generated, pressure of dry hydrogen, temperature, and volume. If necessary, convert volumes from milliliters to liters.



Experiment 12 Report Sheet

Gas Laws II: The Universal Gas Constant

Name _____

Class _____

Date _____

Prelab

1. How will the volume occupied by a gas respond to an increase in the number of moles of gas present, assuming constant temperature and pressure?

_____

2. How will the pressure of a gas respond to a decrease in the number of gas molecules, providing there is no change in temperature or volume?

_____

3. Professional drivers sometimes stop to let some air out of their tires when they are are in the middle of a race during hot weather. Use your knowledge of gas laws to explain this behavior.

_____

_____

4. The accepted value for the universal gas constant is 8.315 kPa•L/mol•K. What would the value be if pressure is measured in torrs and volume is in milliliters?

_____ torr•mL/mol•K

Data and Observations

Summary Table

	Trial 1	Trial 2	Trial 3
number of moles of hydrogen			
pressure of dry hydrogen			
temperature (kelvins)			
volume of hydrogen			
value for *PV/nT*			

Average value for *PV/nT* _____

▶ Calculations

Show all calculations for one of your three trials; you may simply report the results for the other two.

1. Determine the value of *PV/nT* for each trial. Enter the values in the Summary Table and then calculate an average.

2. Your average for *PV/nT* represents your experimental value for the universal gas constant. For pressure in kPa, the accepted value is 8.31 kPa•L/mol•K. Calculate your percent error in the determination of *R*.

Percent error _____ %

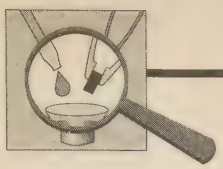
▶ Analysis and Conclusions

1. Imagine that you had to explain Experiments 11 and 12 to someone who had never done them, but whose knowledge of chemistry and laboratory work was about the same as yours. Write a paragraph that explains the goals and the results of these experiments.

2. Challenge: The premise on which the ideal gas law is based is the assumption that gas molecules do not exert any attractive force on each other. However, all gas molecules attract each other at least slightly. Thus no real gas behaves exactly according to the ideal gas law. Suggest explanations for the following observations.

   a. Real gases come closest to ideal behavior when the pressure is low and the temperature is high.

   b. Gases are less likely to exhibit ideal behavior near their condensation temperature. (The condensation temperature is to the boiling point of a substance, as its freezing point is to its melting point.)



Experiment 13

Classifying Elements

Chemical elements may be classified on the basis of their physical and chemical properties. Some can be classified as *brittle,* meaning they will break when subjected to stress, or *malleable,* meaning they will bend under stress. Elements can be classified by physical appearance: Does the surface of a solid have a shiny luster or a dull finish? A third basis for classification is an element's ability to conduct an electric current. All three methods involve physical properties—properties that describe an element on its own, independent of other substances. By contrast, chemical properties involve the interaction of pure substances, including elements, with other pure substances. For example, one can classify elements by the way they react with acids.

In this experiment, you will run five different tests on solid samples of eight elements. The elements will be identified by code letters A–H. Three of the tests involve physical properties; two illustrate chemical behavior.

You will study the physical properties first. Seven of the elements are safe to handle if you wash your hands thoroughly before leaving the lab. Element H is unsafe to handle directly. Your teacher will test element H for brittleness and malleability during the prelab discussion. Your teacher will then dispense samples of H so that you can complete the tests on H. Your teacher will also demonstrate how to use the conductivity tester. Record the results of all tests in Tables 1 and 2 on the Report Sheet.

OBJECTIVES

1. ***Test*** and ***compare*** the physical and chemical properties of several elements.
2. ***Classify*** the eight elements into groups with similar properties.
3. ***Identify*** the groups as containing metals, nonmetals, or metalloids.

▶ Materials

Apparatus

24-well test plate
forceps or tweezers
wash bottle
safety goggles
laboratory apron
gloves

Reagents

0.5 M hydrochloric acid
0.5 M copper(II) chloride solution
samples of elements A–H

Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–3 on the Report Sheet.

Procedure

Part 1: Tests of Physical Properties

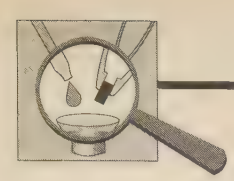
1. Put on your laboratory apron, safety goggles, and gloves.
2. Use forceps or tweezers to place a small piece of each element, A through G, in separate wells of your test plate. Record the appearance of each element—its color and whether its surface is dull or shiny.
3. Using tweezers or forceps, try bending each sample. Does it bend or does it snap or crumble into pieces? If you cannot bend or break the sample, remove it from the well. Fold the sample in a paper towel and place it on a clean, hard surface (such as the base of a ring stand). Then hit it gently with an iron ring or other small instrument. Does pounding cause the sample to flatten or to shatter?
4. Following your teacher's instructions, test each element for conductivity.

Part 2: Tests of Chemical Properties

5. Divide your samples so that you have two wells containing element A, two with element B, and so on. These tests can be done with very small samples; you need only observe whether changes are taking place. (If you need to confirm an observation, obtain an additional small sample.)
6. To one of the wells containing element A, add enough 0.5 M hydrochloric acid, HCl(aq), to just cover the solid—you do not need to fill the well. Look for evidence of chemical reaction such as a color change or evolution of a gas. In the same manner, test the other sample of A with copper(II) chloride solution, $CuCl_2$(aq). Again, look for any sign of chemical change.
7. Repeat steps 5 and 6 for elements B–H.

Cleaning Up

Use a wash bottle to rinse the metal samples and solutions from your test plate onto paper towels in the container labeled *Experiment 13 Waste.* You may need to use tweezers to remove some of the solid pieces. Wash the test plate thoroughly with soap and water. If a well appears stained after washing, notify your teacher. Return all equipment to its proper location and wash your hands thoroughly before leaving the lab.



Experiment 13 Report Sheet

Classifying Elements

Name _____

Class _____

Date _____

Prelab Questions

1. In what order did Mendeleev originally arrange the elements? Why did he make changes in the sequence?

2. An iron coat hanger can be easily bent, but a wrought-iron fence would not be classified as flexible. Discuss how these observations pertain to the flexibility tests you will perform in this experiment.

3. In Part 2, you will test the reactivity of the samples with 0.5 M solutions of copper(II) chloride and hydrochloric acid. These solutions are mainly water. Suppose you detect a color change. How can you be certain that the change is evidence of a reaction and not the result of the element dissolving in water?

Data and Observations

Table 1: Tests of Physical Properties

Element	Malleable or Brittle	Conductivity (Y/N)	Surface (shiny or dull)	Color
A				
B				
C				
D				
E				
F				
G				
H				

Table 2: Tests of Chemical Properties

Element	Reaction with $CuCl_2$	Reaction with HCl
A		
B		
C		
D		
E		
F		
G		
H		

▶ Analysis and Conclusions

1. For each of the five tests of physical and chemical properties, separate the eight elements according to their behavior on that test.

2. Combine your five lists from question 1 to make two groups of elements so that all the members of a group are alike in most of the properties tested. You may find a few elements that have some properties that fit one group and some properties that fit the other. If so, place them in a third category.

3. As you learned in Chapter 2, *metals* show the following characteristics: They are shiny, often gray or silver in color; they are malleable; and they are good conductors of heat and electricity. In addition, they may react with dilute acids to release hydrogen gas. You may recall from Chapter 3 that copper is less active than most other metals. The copper ions in a solution of copper(II) chloride can undergo single replacement reactions with metals that are more active than copper.

Nonmetals tend to be brittle, with a dull or dark finish. They are not good conductors of heat or electricity. Nonmetals do not react with dilute acids or with low-activity cations such as copper(II) ions.

There is a third category called *metalloids*. These elements show some similarities to both metals and nonmetals. They are usually good conductors, but are brittle. They are unreactive to dilute acids and low-activity cations. They may or may not have a shiny surface.

Use your groupings to identify elements A–H as *metals, nonmetals,* or *metalloids.*

Experiment 14

Activity Series I: Metals

In Experiment 13 you classified elements as metals, nonmetals, and metalloids on the basis of their physical and chemical properties. You discovered that there were differences among elements within the same category. For example, zinc and aluminum react with hydrochloric acid, but copper does not. In other words, metals differ in their reactivity, or *activity.*

Metals can form compounds by losing electrons and forming positive ions, or cations. In return, cations can gain electrons and become neutral atoms. As you learned in Chapter 11, some metals lose electrons easily (those with lower ionization energies), while others do so only with difficulty. The relative activity of a metal is related to its ability to lose electrons from its outermost energy level.

When atoms of an active metal are brought in contact with cations of a less active metal, one or more electrons transfer from the active metal to the cations. Thus, for active metal X and less active metal Y, reactions like the following are possible.

$X + Y^{+} \longrightarrow X^{+} + Y$ (one electron transferred)
$X + Y^{2+} \longrightarrow X^{2+} + Y$ (two electrons transferred)

However, cations of the more active metal will not react with atoms of the less active metal.

$Y + X^{+} \longrightarrow$ N.R.

To summarize, ions of less active metals can remove electrons from neutral atoms of more active metals.

In this experiment, you will determine the relative activities of several metals: copper, iron, magnesium, silver, tin, and zinc. You will place samples of the metals in solutions containing the cation form of each of the other metals. Because silver is expensive, your teacher will demonstrate its activity following the experiment.

OBJECTIVES

1. ***Test*** the relative activities of several metals.
2. ***Order*** the metals from most active to least active.

Materials

Apparatus

24-well test plate
forceps or tweezers
wash bottle
laboratory apron
safety goggles
gloves

Reagents

thin-stem pipets containing
- 0.2 M $CuCl_2(aq)$
- 0.2 M $FeCl_3(aq)$
- 0.2 M $Mg(NO_3)_2(aq)$
- 0.2 M $SnCl_2(aq)$
- 0.2 M $Zn(NO_3)_2(aq)$

small pieces of Cu, Fe, Mg, Sn, Zn

Prelab

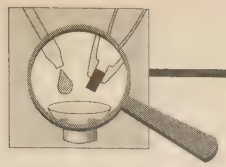
1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–3 on the Report Sheet.

Procedure

1. Put on your safety goggles, laboratory apron, and gloves.
2. Place single pieces of one metal in four different wells of your 24-well test plate. Cover each of the pieces with 8–10 drops of a different metal solution. *Do not use the solution that contains cations of the metal being tested.* Observe carefully because the only sign of reaction may be a darkening of the metal surface. Some reactions will not occur immediately; wait at least 8–10 minutes before you decide that nothing happened. Record evidence of reaction in the Data Table on the Report Sheet. If no reaction occurs, record "N.R."
3. Repeat step 2 for the four other metals. If you are uncertain about the results of any pairing, use one of the empty wells in your test plate to test the metal in question with distilled water as a control.
4. Your instructor may let you leave your well plate overnight in a safe place to see if any further changes occur. If so, label it and obtain a cover.

Cleaning Up

Shake the contents of your test plate into the large pan lined with paper towels. Use forceps or tweezers to remove pieces of metal that remain in the wells. Then use a wash bottle to rinse solutions from the test plate onto the paper towels. Your teacher will complete the disposal process. Before leaving the lab, clean up all other materials and wash your hands thoroughly.



Experiment 14 Report Sheet

Activity Series I: Metals

Name _____

Class _____

Date _____

Prelab Questions

1. Aluminum metal will react with a solution of copper(II) ions, but copper metal will not react with a solution of aluminum ions. Which metal is more active, copper or aluminum?

_____

2. Write a balanced equation for the reaction between aluminum metal and copper(II) chloride.

_____

Data and Observations

Data Table: Test Results

Metals	Cations				
	Zn^{2+}	Sn^{2+}	Mg^{2+}	Fe^{3+}	Cu^{2+}
Cu					
Fe					
Mg					
Sn					
Zn					

Observations

Demonstration of copper + silver ions:

Analysis and Conclusions

1. Which metal reacted with the most solutions containing other metal cations? (It is the most active.)

_____

2. Which metal reacted with the fewest solutions containing other metal cations? (It is the least active.)

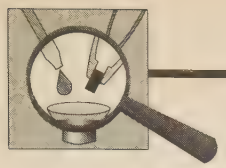
_____

3. Rank the five metals in order of decreasing activity (most active first).

_____

4. Your teacher will place copper metal in a solution containing silver ions. Record your observations. Then assign silver a position on your activity list.

5. Write a balanced equation for each reaction that occurred.



Experiment 15

Activity Series II: Nonmetals

When halogens react with metals to form compounds, they gain electrons and form anions that are isoelectronic with the nearest noble gas. The greater the tendency of a nonmetal to acquire electrons, the more active the nonmetal. In this experiment, you will place halogen molecules (Cl_2, Br_2, and I_2) in direct competition for electrons with halide ions (F^-, Cl^-, Br^-, and I^-). From your results, you should be able to infer the relative activities for the halogen family.

To gather your data, you will take advantage of differences in solubility between halogen molecules and halide ions in water and a liquid called methyl tertiary-butyl ether, or MTBE for short. MTBE is added to unleaded gasoline to boost the octane rating. When you try to mix water and MTBE, they form two distinct layers. Halide ions are very soluble in water, but only slightly soluble in MTBE. By contrast, halogen molecules are soluble in MTBE, but only slightly soluble in water.

OBJECTIVES

1. ***Compare*** the solubilities of halogens in water and MTBE.
2. ***Test*** the reactivities of selected molecular halogens with halide ions.
3. ***Determine*** the relative activity of the four halogens.

▶ Materials

Apparatus

75-mm or 100-mm reaction tubes

24-well test plate or other rack for reaction tubes

corks for reaction tubes

laboratory apron

safety goggles

Reagents

methyl tertiary-butyl ether (MTBE)

thin-stem pipets containing

- iodine water, $I_2(aq)$
- bromine water, $Br_2(aq)$
- chlorine water, $Cl_2(aq)$
- 0.2 M sodium fluoride solution, NaF(aq)
- 0.2 M sodium chloride solution, NaCl(aq)
- 0.2 M sodium bromide solution, NaBr(aq)
- 0.2 M sodium iodide solution, NaI(aq)

sodium thiosulfate solution, $Na_2S_2O_3(aq)$, for disposal of halogens

Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–3 on the Report Sheet.

Procedure

Part 1: Preliminary Tests

1. Put on your laboratory apron and safety goggles.
2. Place 0.5 mL of MTBE in each of three reaction tubes. To one of the tubes add 0.5 mL of chlorine water, $Cl_2(aq)$; add 0.5 mL of bromine water, $Br_2(aq)$, to the second; and place 0.5 mL of iodine water, $I_2(aq)$, in the third. Cork each tube and shake briefly to mix. Record the appearance of the contents of each tube in the Data Table on the Report Sheet. Any color in the MTBE layer is due to the halogen molecule. (When you need to reuse a reaction tube, follow the procedure described under Cleaning Up.)

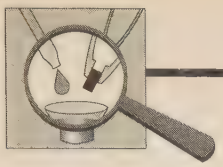
Caution: MTBE is highly flammable. It must be kept away from open flames, spark sources, and hot surfaces.

Part 2: Activity Tests

3. To keep track of the reagents in Part 2, label the reaction tubes *NaF, NaBr, NaCl,* and *NaI,* respectively. Place 0.5 mL of NaF(aq), NaBr(aq), and NaI(aq) in the appropriate reaction tube. Add 0.5 mL of MTBE to each tube. Then add 0.25 mL of chlorine water to each tube. Cork the tubes and shake briefly to mix. Allow the contents to settle and record the appearance of both layers in each tube.
4. Repeat step 3 with 0.5 mL of NaF(aq), NaCl(aq), and NaI(aq), substituting bromine water for chlorine water.
5. Repeat step 3 with 0.5 mL of NaF(aq), NaCl(aq), and NaBr(aq), substituting iodine water for chlorine water.

Cleaning Up

Both MTBE and the molecular halogens must be disposed of with caution. Add about 10 drops of saturated sodium thiosulfate solution, $Na_2S_2O_3(aq)$, to each reaction tube. Cork the tube and shake gently to convert the halogens to halide ions. If the color of the halogen remains in the MTBE layer, add 5 additional drops, cork, and shake again. Dispose of the contents of the tubes in the container supplied by your instructor. Before leaving the lab, clean up all materials and wash your hands thoroughly.



Experiment 15 Report Sheet

Activity Series II: Nonmetals

Name _____

Class _____

Date _____

Prelab Questions

1. Write the electron configuration for each neutral halogen atom and for the corresponding halide ion.

_____

_____

_____

2. Identify the noble gas with which each of the halide ions is isoelectronic.

_____

3. Water and MTBE form separate layers when mixed. What physical property determines which liquid will form the top layer and which the bottom?

_____

Data and Observations

Data Table

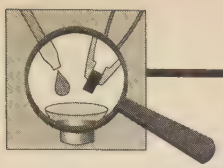
Part 1: Preliminary Tests	
MTBE + Cl_2(aq)	
MTBE + Br_2(aq)	
MTBE + I_2(aq)	
Part 2: Activity Tests	
Cl_2(aq) + NaF(aq)	
Cl_2(aq) + NaBr(aq)	
Cl_2(aq) + NaI(aq)	
Br_2(aq) + NaF(aq)	
Br_2(aq) + NaCl(aq)	
Br_2(aq) + NaI(aq)	
I_2(aq) + NaF(aq)	
I_2(aq) + NaCl(aq)	
I_2(aq) + NaBr(aq)	

▶ Analysis and Conclusions

1. Write equations for any reactions that occurred.

2. Examine your data and place bromine, chlorine, and iodine in order of relative activity (most active first). Remember that a more active halogen will remove electrons from the ions of a less active halogen. Based on your activity list and the positions of the halogens on the periodic table, determine the appropriate position for fluorine on your list.

3. If you were able to mix $F_2(aq)$ with solutions of NaCl, NaBr, and NaI, what reactions would your expect?



Experiment 16

Dyes and Dyeing

Cloth is made from fibers, which consist of millions of long-chain molecules. Some natural fibers, such as silk and wool, are composed of protein molecules. Proteins contain both polar and ionic bonding sites. Because most dye molecules are either polar or ionic, they bond readily to a fabric such as wool. In contrast, dyes bond less readily to fabrics such as nylon, whose fibers have fewer polar sites. Fabrics such as Dacron™ and Orlon™, with scattered polar sites along their molecular chains, show intermediate attraction for dye molecules. In this experiment, you will use intensity of color as an indicator of the attraction of a fabric for dye—the deeper the color, the stronger the attraction. The fabric samples that you will use contain thirteen different fibers arranged in an ordered pattern.

OBJECTIVES

1. ***Compare*** the dye-absorbing ability of different fabrics.
2. ***Test*** the colorfastness of the dyed fabrics.
3. ***Identify*** the fabrics based on test results.

▶ Materials

Apparatus

forceps or tweezers
paper towels
2 small zipper-type plastic bags
warm water bath
safety goggles
laboratory apron
gloves

Reagents

multifiber fabric test strips
malachite green
methyl orange
Congo red
indigo
Test Dye #3
vinegar

▶ Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–3 on the Report Sheet.

▶ Procedure

Put on your laboratory apron, gloves, and safety goggles.

Part 1: Direct Dyeing

Because the dyes malachite green and methyl orange have both ionic and polar bonding sites, they can attach to similar sites on the fabric fibers. One partner will test methyl orange; one will test malachite green.

Caution: Dyes will stain skin as well as fabric. Some dyes can also irritate your skin.

Caution: Contact with simmering aqueous solutions can produce serious burns. If some hot dye should splash on you or your clothing, wash thoroughly with plenty of cool running water.

Caution: Do not wash dyed strips with other items to be laundered.

Immerse a test strip in the simmering dye bath for 5–10 minutes. Use forceps or tweezers to remove the strip from the bath; allow as much of the dye solution as possible to drain back into the bath. Then rinse the fabric with cool running tap water to remove excess dye. Place the strip on a labeled paper towel to dry. Wrap the dried strip in the paper towel and place it in a plastic bag to carry home.

At home, test the strip for colorfastness. Cut it in half lengthwise and hand wash one of the half-strips with soap and water. Compare the color intensity of the washed and unwashed strips.

Part 2: Substantive Dyes

Direct dyeing works best with fabrics that have both ionic and polar bonding sites. Fabrics such as cotton and rayon, which have only polar sites, require a different type of dye. **Substantive dyes** are large colloidal molecules that can attach to cotton and rayon by hydrogen bonds. One partner will dye a test strip with the substantive dye Congo red, while the other partner completes Part 3.

Immerse a 2-cm wide test strip in the Congo red bath for 5–10 minutes. Use forceps or tweezers to remove the strip from the bath. Drain off excess dye solution and place the fabric strip in a warm water bath (about 40°C). Wash in the warm water until all excess dye is removed. Wring out excess water, and then cut the strip in half lengthwise. Dip the half-strips in white vinegar and observe the result. Rinse one of the two half-strips in tap water. Place them on labeled paper towels to dry. Wrap the dried strips in their respective paper towels and place them in your plastic bag. At home, hand wash both strips with soap to test their colorfastness.

Part 3: Indigo—a Vat Dye

Indigo is used to dye denim (GUESS™ what color it is!). Because indigo is not very soluble in water, it does not rinse out very easily. However, indigo does rub off easily because it is bound to the fabric by relatively weak intermolecular forces, primarily hydrogen bonds. Indigo can be removed by bleach, which reacts with the dye to form a more soluble product.

Immerse a test strip in the indigo bath and boil gently for about 5–10 minutes. Use forceps or tweezers to remove the strip from the bath and rinse it well in cool tap water. Let it dry on a labeled paper towel. Wrap the dried strip in the paper towel and place it in your plastic bag.

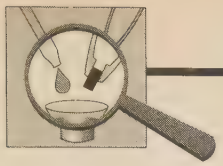
Caution: When using bleach, follow the instructions and wear rubber gloves.

At home, cut the strip in half lengthwise. Hand wash one half in mild soap solution, and the other in the same soap solution with a little bleach added. (For a bathroom sink, less than a capful of liquid bleach is sufficient.) Compare the color intensity of the two half-strips.

Part 4: Fabric Analysis

You will test a dye mixture called Test Dye #3 to identify the fibers in the test strips. Soak a 1-cm swatch in Test Dye #3 for 5–10 minutes. Use forceps or tweezers to remove the swatch from the dye and drain off any excess dye solution. Dry the swatch on a labeled paper towel. One partner should take the strip home to test for colorfastness.

▶ ***Cleaning Up*** Your teacher will dilute the hot dye baths and flush the dyes down the sink. Clean up all other equipment and wash your hands thoroughly before leaving the lab.



Experiment 16 Report Sheet

Dyes and Dyeing

Name _____

Class _____

Date _____

Prelab Questions

1. Draw the electron dot structure for methyl alcohol (CH_3OH) and methyl amine (CH_3NH_2). Hint: On methyl amine, start with the bond between the carbon atom and the nitrogen atom.

2. Refer to Figures 14-5 and 14-6 on page 475 of your text. Would you expect the C–O–H bond in methyl alcohol to be linear? Explain.

_____

3. The cellulose molecules in cotton contain many C–O–H groups similar to those in methyl alcohol. The proteins in wool contain NH_2 groups attached to carbon atoms as in methyl amine. Will cotton and wool tend to form hydrogen bonds?

_____

_____

Data and Observations

Part 1: Direct Dyeing

Part 2: Substantive Dyes

Part 3: Indigo—a Vat Dye

▶ Analysis and Conclusions

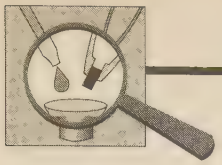
Because each partner collected different data, prepare a single report for the team. Fasten the labeled test strips to the report to illustrate your answers.

1. Compare the swatch that was tested with Test Dye #3 with the color coded strip that your teacher will display. Identify the fabrics. What marker will you use to identify the fabrics in the other samples?

2. Of the dyes tested, which seems to be the most colorfast? Which of the individual fibers are the best at retaining color?

3. Of the fibers tested, which one appears to have the fewest polar bonding sites? Provide evidence to support your conclusion.

4. Compare the results with fabrics 6 and 7, Dacron 54 and Dacron 64. Dacron 64 contains an optical brightener. Based on your data, how does an optical brightener affect the intensity of the color?



Experiment 17

Vapor Pressure of Water

When a liquid evaporates in a closed container, vapor accumulates above the liquid. The partial pressure of the vapor is called the **vapor pressure.** The vapor pressure depends on the number of gas particles. The more vapor accumulates, the greater the vapor pressure. An increase in temperature increases the rate of evaporation and the vapor pressure.

OBJECTIVES

1. ***Measure*** the volume of a sample of moist air at several temperatures.
2. ***Calculate*** the vapor pressure of water at each temperature.
3. ***Graph*** vapor pressure as a function of temperature.

▶ Materials

Apparatus

10-mL graduated cylinder (without plastic base)
tall-form 1-L beaker or similar heat-tolerant container
hot plate
thermometer, –20°C to 110°C
long stirring rod
beaker tongs or heat-proof mitt
large tray or sink
ice
safety goggles

Reagents

tap water

▶ Prelab

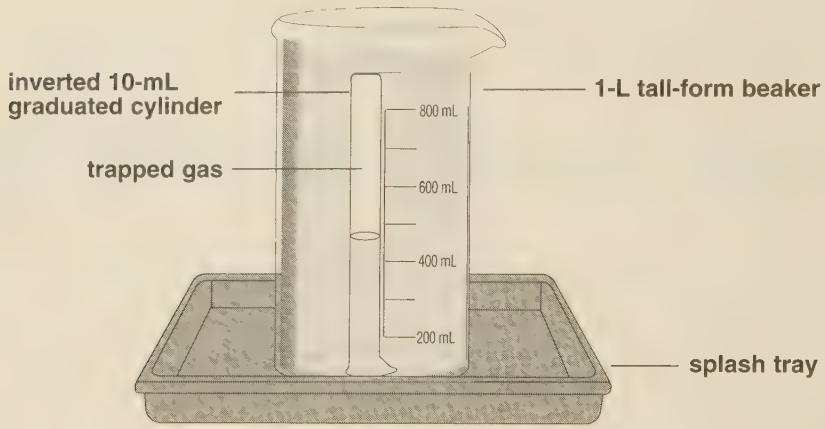
1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–3 on the Report Sheet.

▶ Procedure

1. Put on your safety goggles.
2. Place 7–8 mL of water in a 10-mL graduated cylinder. Using your fingers as a plug, invert the graduate in a tall, heat-tolerant container of water. The *entire* graduated cylinder must be under water. Heat the container on a hot plate to a temperature of about 75°C. Using a long stirring rod, gently stir the water to ensure even heating. Use the rod to gently dislodge air bubbles from the wall of the cylinder. One partner should obtain and record the atmospheric pressure.
3. Use beaker tongs or a heat-proof mitt to remove the container from the heat source. Read the volume of gas in the cylinder to the nearest 0.1 mL. Read the temperature of the water in the container at about the midpoint of the column of gas. Record the water and volume values in the Data Table on the Report Sheet.

Caution: 75°C water can cause a serious burn; rinse the affected area with cool running water. *Do not* use the thermometer as a stirring rod.

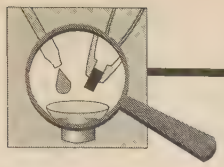
4. Place the system in a tray that can catch overflowing water as shown in the illustration below. Add ice, one or two pieces at a time, stirring gently to maintain a uniform temperature throughout the system. Allow each portion of ice to melt before adding more. Record the volume of gas in the graduated cylinder at 5-degree intervals until the temperature reaches 40°C. Record the volumes and temperatures in the Data Table.



5. Add handfuls of ice while stirring gently until the water cools to below 5°C. Then record the volume and temperature again.

▶ Cleaning Up

Pour the water down the sink. Return all equipment to the proper locations and wash your hands before leaving the laboratory.



Experiment 17 Report Sheet

Vapor Pressure of Water

Name _____

Class _____

Date _____

Prelab Questions

1. Write the ideal gas law (a) in standard form; (b) rearranged to solve for n; and (c) rearranged to solve for P.

2. At any given temperature, what gas other than air will be in the inverted cylinder?

3. State Dalton's Law of Partial Pressures and explain how it applies to this experimental procedure.

Data and Observations

Data Table

Barometric Pressure (torr): _____

$V_{observed}$ (mL)	$V_{corrected}$ (mL)	T (K)	n_{air} (mol)	P_{air} (torr)	P_{vapor} (torr)

Calculations

1. The graduated cylinder was calibrated to be read in an upright position. Because your readings were made with the cylinder inverted, the volume must be corrected by subtracting 0.2 mL from each observed value. Record the corrected volumes in the Data Table.

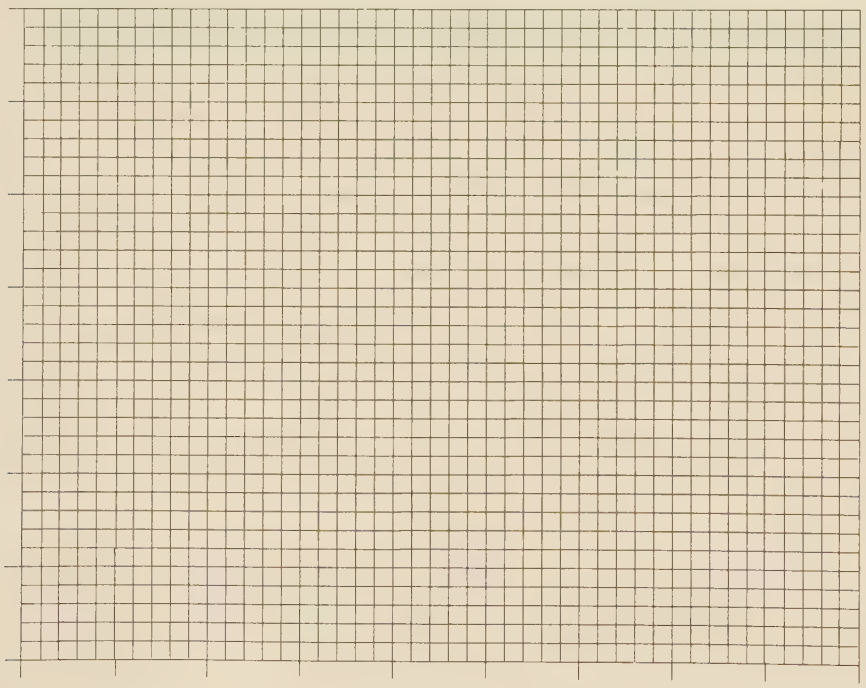
2. Because the vapor pressure of water is very small below 5°C, the gas in the graduated cylinder is assumed to be only air. Thus, we can also assume that the pressure in the cylinder is equal to the barometric pressure. Use these assumptions and the ideal gas law to calculate the number of moles, *n,* of air in the graduated cylinder.

3. Use your corrected volume measurements and the value of *n* obtained in question 2 to calculate the pressure of the trapped air at each of the other temperatures. Record your results in the Data Table. (Your answers should be *less than* the barometric pressure, because at each of these temperatures the graduate contains both air and water vapor.)

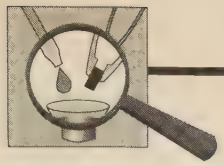
4. At each temperature, the difference between the air pressure and the barometric pressure is the vapor pressure of water at that temperature. For each temperature, calculate the vapor pressure and enter the values in the Data Table.

▶ Analysis and Conclusions

1. Use the grid to plot vapor pressure as a function of Celsius temperature. Draw the best smooth-curve fit you can.



2. Compare your results with the accepted values and curve shape, as found in your text. Discuss your findings.



Experiment 18

Double Replacement Reactions in Solution

In this experiment, you will test ten different solutions in pairs, looking for the formation of a precipitate as evidence of reaction. On a microscale level, precipitation often amounts to little more than a faint cloudiness, which is best detected by looking down through your test plate well at a line of type. Some of the pairs will react; some will not. For those pairs that do react, you will write balanced equations and identify the precipitates by name and formula.

OBJECTIVES

1. ***Test*** each pair of reagents for evidence of reaction.
2. ***Identify*** the precipitates that form.
3. ***Write*** balanced molecular, ionic, and net ionic equations for all reactions.

▶ Materials

Apparatus

96-well test plate (for performing tests)
24-well test plate (for holding reagents)
plastic toothpick stirrer
wash bottle (for cleaning 96-well plate)
cotton swab
safety goggles
laboratory apron
gloves

Reagents (in microtip pipets)

Group A	Group B
sodium iodide, NaI	cobalt(II) chloride, $CoCl_2$
sodium carbonate, Na_2CO_3	copper(II) chloride, $CuCl_2$
sodium phosphate, Na_3PO_4	aluminum chloride, $AlCl_3$
sodium sulfate, Na_2SO_4	strontium chloride, $SrCl_2$
sodium hydroxide, NaOH	nickel(II) chloride, $NiCl_2$

Prelab

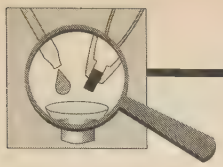
1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–4 on the Report Sheet.

Procedure

1. Put on your laboratory apron, safety goggles, and gloves.
2. Use 3–4 drops of each reagent for each test. Note and record any evidence of reaction. Test all possible pair combinations of each Group A reagent with all Group B reagents. With the Data Table on the Report Sheet as a guide, devise a systematic approach that minimizes the manipulation of pipets.
3. Repeat a test if you are uncertain of the result.
4. In an unused well, place 3 drops of aluminum chloride solution. Add 15 drops of sodium hydroxide one drop at a time. Observe and record what happens with each addition.

Cleaning Up

Return the pipets containing unused reagents to the proper location. Because some of the solutions contain transition and heavy metal ions, they should not be washed down the drain. Instead, dispose of the contents of your well plate onto paper towels in the tray labeled *Transition Metal Waste*. Then rinse your well plate, use a cotton swab to remove any residue from the wells, and rinse again. Discard the swabs in the *Transition Metal Waste* tray. Wash your hands thoroughly before leaving the laboratory.



Experiment 18 Report Sheet

Double Replacement Reactions in Solution

Name _____

Class _____

Date _____

Prelab Questions

1. Write the formulas for the following ionic compounds.

a. zinc sulfide

b. chromium(III) hydroxide

c. lead(II) phosphate

2. Write equations for the dissociation of the following ionic compounds in water.

a. sodium chloride

_____

b. copper(II) chloride

_____

c. iron(III) sulfate

_____

3. Write molecular, ionic, and net ionic equations for the following reactions.

a. iron(III) nitrate reacts with potassium carbonate, producing iron(III) carbonate precipitate and aqueous potassium nitrate.

_____

_____

_____

b. When aqueous solutions of strontium chloride and silver sulfate are mixed, two precipitates are formed.

_____

_____

_____

4. Equal volumes of 0.2 M sodium chloride and 0.2 M silver nitrate are mixed. What will be the molar concentration of each of these ions?

_____

▶ Data and Observations

Data Table: Evidence of Reaction

	NaI	Na_2CO_3	Na_3PO_4	Na_2SO_4	NaOH
$CoCl_2$					
$CuCl_2$					
$AlCl_3$					
$SrCl_2$					
$NiCl_2$					

Observations of drop-by-drop mixing of $AlCl_3$ and NaOH (step 4):

▶ Analysis and Conclusions

1. For each reaction in which a precipitate formed, write the following:

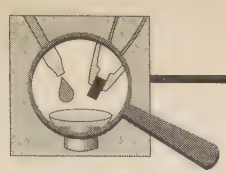
a. the formula for the precipitate

b. a balanced molecular equation

c. a balanced ionic equation, showing all dissolved species as ions and all precipitates as molecules

d. a balanced net ionic equation

2. If there were any combinations that appeared to react, but in which no precipitate formed, try to identify the ions responsible for the reaction. (Hint: Which of the ions were spectator ions in other reactions?)

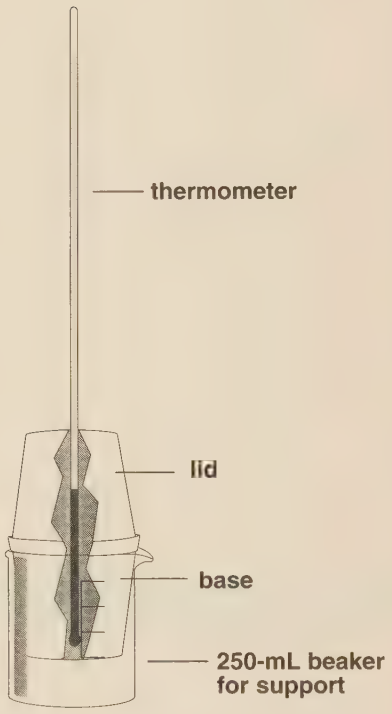


Experiment 19

Calorimetry

In this experiment, you will determine the heat of fusion of ice—the quantity of heat needed to melt one gram of ice. You will use a *calorimeter,* a device that measures the heat associated with a chemical or physical change. The calorimeter is insulated to minimize heat exchange between the system and its surroundings. However, the calorimeter itself can absorb heat, so you must first determine its heat capacity. Heat capacity (C), measured in Joules per degree, compares the heat absorbed by a substance to the resulting increase in temperature.

The calorimeter you will use is shown in the margin. To illustrate the importance of calorimeter design, one partner will use the calorimeter as shown (closed) and one will use the same device, but without the lid (open). Both partners will complete a set of experiments and then compare results.



Calorimeter with lid

OBJECTIVES

1. ***Determine*** the heat capacity of a calorimeter.
2. ***Explain*** the effect of calorimeter design on the experimentally-determined heat capacity.
3. ***Determine*** the heat of fusion of ice.

▶ Materials

Apparatus

3 plastic foam cups
2 250-mL beakers to support calorimeters
4 thermometers, –20°C to 110°C
3 25-mL graduated cylinders
2 thin-stem pipets
milligram or centigram balance
paper towels
safety goggles
laboratory apron

Reagents

tap water
ice

▶ Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–4 on the Report Sheet.

Procedure

1. Put on your safety goggles and laboratory apron.
2. Decide which partner will use the closed calorimeter and which will use the open calorimeter. The one who uses the closed calorimeter must make all measurements with the lid in place and the thermometer protruding through an opening in the lid. It is essential that the thermometer fit tightly in the opening. The partner using the open calorimeter will ignore all references to the lid in the instructions. Record all data in Table 1 on the Report Sheet.

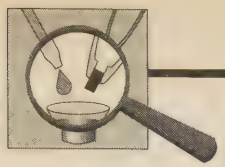
Part 1: The Heat Capacity of the Calorimeter.

3. Place precisely 25.0 mL of warm water in a 25-mL graduated cylinder. (If necessary, use a pipet to adjust the volume.) Insert a thermometer into the graduated cylinder. Place the cylinder with thermometer into one of the communal warm water baths for at least 5 minutes to allow the two water temperatures to equalize. (You may need to remove the plastic base from the graduate to fit it in the water bath.)
4. Pour 25.0 mL of chilled water into the calorimeter. After a few minutes, the calorimeter and the water should be at the same temperature. Just before doing step 5, record the precise temperature of the water.
5. Collect the graduate from the hot water bath. Record the temperature of the water in the graduate. *Immediately* pour all the hot water from the graduate into the calorimeter and replace the lid. It is critical that there be no delay in transferring the hot water because hot water cools rapidly in air. Record the highest temperature reached after mixing.
6. Dry out the calorimeter cup and repeat steps 3–5. If the data from your two trials differ greatly, consult with your teacher to determine whether you need to do a third trial.

Part 2: Determining the Heat of Fusion of Ice.

7. Determine the mass of the empty, dry calorimeter base (without the lid, thermometer, and support beaker). Add about 50 mL of water. Measure the mass of the cup and water to determine the mass of the water. Replace the lid and thermometer.
8. Wait a minute for the temperature to stabilize; then measure and record the temperature.
9. Blot a small piece of ice (about 2–3 cm^3) with a paper towel to remove melted water and add the ice to the calorimeter. Quickly cover the calorimeter and gently swirl the contents. Record the lowest temperature reached. After the temperature begins to rise, remove the lid and thermometer. Measure the mass of the cup and contents to determine the mass of ice that melted.
10. Do at least two trials—three if you have the time. Dry the inside of the calorimeter with a paper towel between trials and use fresh water for each trial.

Cleaning Up Pour the water down the drain. Return all apparatus to the appropriate locations.



Experiment 19 Report Sheet

Calorimetry

Name _____

Class _____

Date _____

▶ Prelab Questions

1. What quantity of heat is absorbed by 100. grams of water if the temperature of the water rises from 15.0°C to 50.0°C? Express your answer in both joules and kilojoules.

2. When you work with recorded temperatures, you know that you must add 273.15 to convert temperatures in degrees Celsius to kelvins. Explain why temperature conversions are not an issue in a calorimetry experiment.

_____

_____

3. A 20.0-g piece of magnesium and a 20.0-g piece of iron are both heated to 100.0°C. They are quickly transferred to separate plastic foam cups, each containing 100.0 g of water at 25°C. With iron, the water temperature rises to 26.6°C; with magnesium, the water temperature rises to 28.5°C. Explain the observed difference in ΔT.

_____

_____

4. Why is it important to use the same calorimeters for Parts 1 and 2 of the procedure?

_____

▶ Data and Observations

Table 1

Type of Calorimeter (circle) open closed			
Part 1	**Trial 1**	**Trial 2**	**Trial 3**
Initial temperature of cool water (step 4)			
Initial temperature of warm water (step 5)			
Highest temperature reached (step 5)			
Part 2	**Trial 1**	**Trial 2**	**Trial 3**
Mass of empty calorimeter cup (step 7)			
Mass of calorimeter and water (step 7)			
Mass of water (step 7)			
Initial temperature of water (step 8)			
Lowest temperature reached (step 9)			
Mass of calorimeter and contents (step 9)			
Mass of ice that melted (step 9)			

Calculations

Part 1: The Heat Capacity of the Calorimeter

When you mixed the warm and cool water, heat was transferred from the warm water to the cool water. You might expect the highest temperature reached by the mixture to be halfway between the initial temperatures of the water samples. However, the final temperature will be lower than expected because the calorimeter absorbs some heat. The difference between the heat lost by the hot water and the heat gained by the cold water must be the amount of heat absorbed by the calorimeter.

To determine the calorimeter constant, *C*, complete the following calculations. Enter your answers in Table 2.

1. a. Calculate the temperature change (ΔT) for the hot water.

b. Calculate the temperature change (ΔT) for the cool water.

2. Calculate the apparent heat loss *(q)* for the hot water.
q = mass of water $\times \Delta T \times 4.18$ J/g°C

3. Calculate the apparent heat loss *(q)* for the cool water.

4. Notice that the heat lost is greater than the heat gained. This difference is the heat absorbed by the calorimeter. Calculate the heat gained *(q)* by the calorimeter.

5. The calorimeter constant for your calorimeter would be the heat gained per degree of temperature change. Calculate the constant, *C*, by dividing q for your calorimeter (answer to question 4) by (ΔT) for the cool water (answer to question 1b).

Table 2

Type of Calorimeter (circle)	open	closed	
	Trial 1	**Trial 2**	**Trial 3**
ΔT for hot water			
ΔT for cool water			
Apparent heat loss for hot water			
Apparent heat loss for cool water			
Heat gained by the calorimeter			
Calorimeter constant			

Part 2: Determining the Heat of Fusion of Ice

6. Calculate the temperature change for the water.

7. Calculate the amount of heat absorbed by the melting ice from the water.

8. The calorimeter is also a source of heat. It releases the same quantity of heat for a decrease in temperature that it absorbs for an equivalent increase in temperature. To calculate the heat absorbed by the melting ice from the calorimeter, multiply the calorimeter constant by the temperature change.

9. Add the results of calculations 7 and 8 to get the total quantity of heat absorbed by the melting ice.

10. Finally, calculate the heat of fusion of ice in J/g, and the molar heat of fusion, in kJ/mol. Show your calculations.

11. Compare your experimental value for heat of fusion with the accepted value (Table 16-2 on page 538 of your text) and calculate your percent error.

▶ Analysis and Conclusions

1. How did calorimeter design affect the precision of your results in Parts 1 and 2 and the magnitude of *C*?

2. Is the calibration of the calorimeter a necessary part of the experiment? If you ignored the heat capacity of the calorimeter, how would this affect the experimental heat of fusion and the percent error?

3. Compare your findings with those of other groups. Were results generally consistent from group to group? How consistent were the results of the individual parts of the experiment from one group to the next?

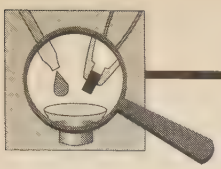
▶ Optional Extensions

If you wish to do one of the following extensions, produce a detailed plan and obtain your teacher's approval. Do not do an unsupervised extension.

1. Determine the specific heat of a sample of metal. Starting with about 50 g of water in the calorimeter, add a solid piece of metal that has been in boiling water for at least 3 minutes.
2. Using 25.0 mL each of 2.0 M HCl(aq) and 2.0 M NaOH(aq), determine the heat of reaction for the following reaction.

   $HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$

   Report your answer in kJ/mol of water and calculate your percent error.
3. Determine the molar heat of solution for solutes such as calcium chloride, potassium nitrate, ammonium and acetone in water. The total volume of liquid in the calorimeter should be about 50 mL—the amount of water you used to find the heat capacity of your calorimeter. Assume that small amounts of solid (about 0.050 mol) will make no significant difference in the volume of water, and that volumes of liquids are additive.
4. Recipes for baking items such as brownies provide different oven temperatures depending on the type of baking pan. For an aluminum pan, the correct temperature is usually 350°C; for a glass pan, it is usually 325°C. Use your knowledge of specific heat to offer an explanation for the difference in recommended oven temperature.



Experiment 20

Investigating Reaction Rates

In this experiment, you will determine the effect of temperature on a reaction rate. You will react cinnamaldehyde, the compound responsible for the distinctive odor of cinnamon, with acetone, the main ingredient in nail polish remover. In the presence of a strong base (KOH), one acetone molecule reacts with a molecule of cinnamaldehyde, forming an intermediate. This intermediate reacts with a second molecule of cinnamaldehyde and forms a yellow solid, which precipitates from the solution. The overall equation for the reaction is:

$$\underset{\text{acetone}}{C_3H_6O} + \underset{\text{cinnemaldehyde}}{2C_9H_8O} \xrightarrow{KOH} \underset{\text{diphenylnonatetrene}}{C_{21}H_{18}O} + \underset{\text{water}}{2H_2O}$$

During the reaction, the yellow color of the solution intensifies, culminating with the sudden formation of the precipitate. Reaction is considered complete when the yellow solid first appears.

OBJECTIVE

1. ***Measure*** time to completion for a reaction between cinnamaldehyde and acetone.
2. ***Determine*** the rate of reaction in moles per liter per second.
3. ***Determine*** the effect of temperature on the rate of the reaction.

▶ Materials

Apparatus

3 10-mL graduated cylinders
3 25-mL Erlenmeyer flasks
stopwatch or clock with sweep second hand
150-mL beaker
600-mL beaker
hot plate (for Part 2)
ice (for Part 3)
thermometer, –20°C to 110°C
safety goggles
laboratory apron
gloves (non-permeable)

Reagents

10% acetone solution by volume in 95% ethanol
2.0 M potassium hydroxide solution
20% cinnamaldehyde solution by volume in 95% ethanol

▶ Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–4 on the Report Sheet.

Procedure

Part 1: Reaction Rate at Room Temperature

1. Put on your safety goggles, laboratory apron, and gloves. Many gloves are *permeable* to one or more of the reagents in this experiment, that is, reagents can penetrate the tiny openings in the gloves. Wearing permeable gloves can provide a false sense of security. Reagents trapped between a glove and your skin can be more harmful than reagents that are spilled on the skin and immediately rinsed off with a stream of cool running water.
2. To ensure that you do not mix reactants in your graduated cylinders, label the cylinders A for acetone, *B* for the base KOH, and *C* for cinnamaldehyde. If you accidentally pour a solution into the wrong cylinder, rinse it with 95% ethanol, wash with soap and water, and dry it thoroughly before proceeding.

Caution: 2 M KOH is corrosive. If it gets into your eyes, use the eyewash fountain immediately. Wash spills or splashes off skin or clothing with a continuous stream of cool tap water and notify your teacher. Cinnamaldehyde can irritate sensitive skin areas, such as the skin on your face.

3. Measure out 5.0 mL of cinnamaldehyde solution, 3.0 mL of 2 M KOH, and 2.0 mL of the acetone solution.
4. Pour the cinnamaldehyde solution into a clean, dry 25-mL Erlenmeyer flask. Add the KOH solution. (Any cloudiness will clear up as soon as you add the acetone, solution *A*.)
5. Add the acetone solution *all at once* and begin timing *immediately.* Gently swirl the flask continuously until the yellow precipitate forms, which marks the completion of reaction. Record the reaction time to completion and the room temperature in the Data Table on the Report Sheet. Before continuing, empty the contents of your flask into the container labeled *Experiment 20 Waste.*
6. Repeat steps 3–5 twice, using a clean, dry flask for each trial.

Part 2: The Effect of Increased Temperature

Your teacher will assign Part 2 to some teams and Part 3 to other teams. You will need to clean your 25-mL flasks. See instructions in Cleaning Up.

Caution: The acetone and cinnamaldehyde solutions are flammable. There should be no open flames in the lab during this procedure.

7. Use a hot plate to warm water in a 600-mL beaker to between 35° and 40°C.
8. Measure out 5.0 mL of cinnamaldehyde solution, 3.0 mL of 2 M KOH, and 2.0 mL of the acetone solution in the appropriate graduated cylinders. Heat the cylinders in a warm water bath as shown in Figure 1 to equalize the temperatures of the solutions and the water. Monitor the temperature of the water bath because acetone is very volatile and will evaporate rapidly at temperatures above 45°C.
9. After 4–5 minutes, pour the cinnamaldehyde solution into a clean, dry 25-mL Erlenmeyer flask. Add the KOH solution. As you remove each graduate from the water bath, give the graduate a quick wipe with a paper towel to prevent water from getting into the Erlenmeyer flask.

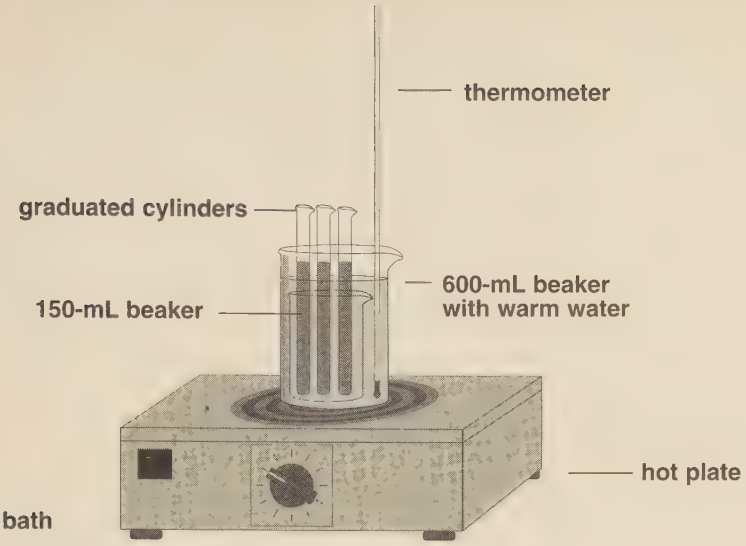


Figure 1: Warm water bath

10. Remove the cylinder containing the acetone from the bath. Wipe it quickly with a paper towel. Then add the acetone solution *all at once* to the Erlenmeyer flask and begin timing *immediately.* To keep the reaction mixture at the same temperature as the water bath, hold the Erlenmeyer flask in the water bath as shown in Figure 2. Gently swirl the flask continuously until the yellow precipitate forms. Record the reaction time to completion and the water bath temperature in the Data Table. Before continuing, empty the contents of your flask into the container labeled *Experiment 20 Waste.*

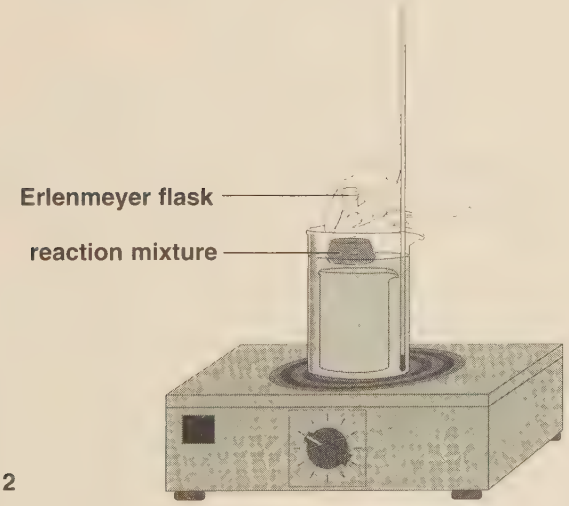


Figure 2

11. Repeat steps 8–10 twice, using a clean, dry flask for each trial. Try to keep the bath at about the same temperature for each trial, but record the precise temperature for each trial, even if the differences are slight.

Part 3: The Effect of Reduced Temperature

12. Use ice to prepare and maintain a cool water bath at a temperature between 5°C and 15°C. At temperatures below 5°C, the reaction is too slow to time accurately.
13. Measure out 5.0 mL of cinnamaldehyde solution, 3.0 mL of 2 M KOH, and 2.0 mL of the acetone solution in the appropriate graduated cylinders. Cool the cylinders in the cool water bath as shown in Figure 3 to equalize the temperatures of the solutions and the water.

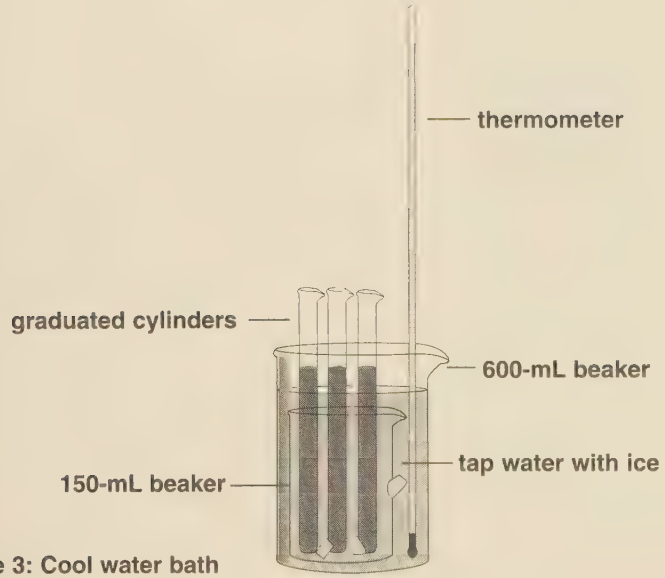


Figure 3: Cool water bath

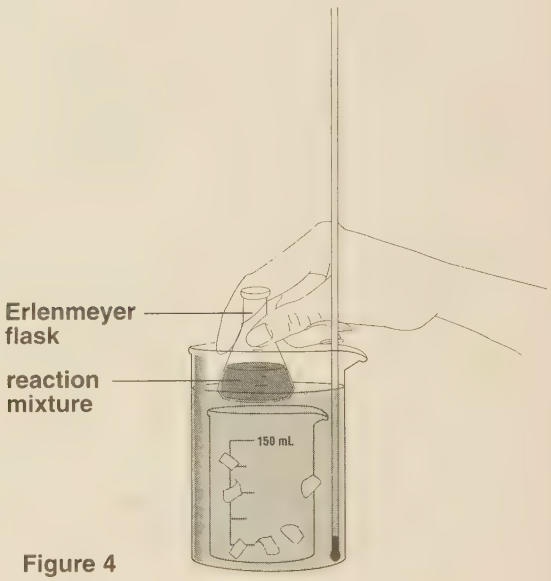
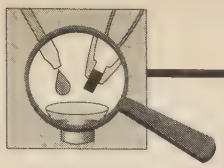


Figure 4

14. After 4–5 minutes, pour the cinnamaldehyde solution into a clean, dry 25-mL Erlenmeyer flask. Add the KOH solution. As you remove each graduate from the water bath, give the graduate a quick wipe with a paper towel to prevent water from getting into the Erlenmeyer flask.
15. Remove the cylinder containing the acetone from the bath. Wipe it quickly with a paper towel. Then add the acetone solution *all at once* to the Erlenmeyer flask and begin timing *immediately.* To keep the reaction mixture at the same temperature as the water bath, hold the Erlenmeyer flask in the water bath as shown in Figure 4. Gently swirl the flask continuously until the yellow precipitate forms. Record the reaction time to completion and the water bath temperature in the Data Table. Before continuing, empty the contents of your flask into the container labeled *Experiment 20 Waste.*
16. Repeat steps 13–15 twice, using a clean, dry flask for each trial. Try to keep the bath at about the same temperature for each trial, but record the precise temperature for each trial, even if the differences are slight.

▶ ***Cleaning Up*** Rinse used glassware with 95% ethanol, then wash with soap and water. Wash your hands thoroughly each time you clean glassware and before leaving the laboratory.



Experiment 20 Report Sheet

Investigating Reaction Rates

Name _____

Class _____

Date _____

▶ Prelab Questions

1. Explain why increasing the temperature usually results in an increase in the rate of a reaction.

_____

_____

_____

2. The cinnamaldehyde solution is 20.0% cinnamaldehyde by volume; that is, 20.0 mL of cinnamaldehyde for each 100. mL of solution. Pure cinnamaldehyde has a density of 1.05 g/mL.

a. What mass of cinnamaldehyde is present in 100. mL of solution? How many moles of cinnamaldehyde are present?

b. Calculate the molarity of the cinnamaldehyde solution.

3. The 10% acetone solution contains 10.0 mL of acetone per 100. mL of solution. Acetone has a density of 0.791 g/mL.

a. Calculate the mass in grams and the number of moles of acetone present in 100. mL of solution.

b. Calculate the molarity of the acetone solution.

4. Given the concentrations of the two reactants and the ratio of volumes that are to be mixed, which is the limiting reactant: cinnamaldehyde or acetone?

Data and Observations

Data Table

Part 1:	Room Temperature: _____°C		
	Trial 1	**Trial 2**	**Trial 3**
Reaction time (s)			
Part 2 or Part 3:	**Trial 1**	**Trial 2**	**Trial 3**
Reaction Time (s)			
Temperature (°C)			

Analysis and Conclusions

Part 1: Reaction Rate at Room Temperature

1. Calculate the average reaction time for your three trials. If the data from one trial is noticeably different, do not use it to find the average.

2. Using your answer to Prelab question 4, calculate the number of moles of acetone consumed in the reaction.

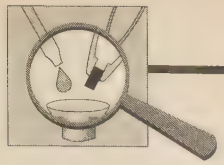
3. Reaction rates are expressed in units of molarity per unit time. Use your answers to questions 1 and 2 to calculate the average reaction rate in moles per liter of acetone consumed per second (M/s).

Part 2 or Part 3: Effect of Temperature on Reaction Rate

4. Calculate the average reaction time for your three trials. If the data from one trial is noticeably different, do not use it to find the average.

5. Use your answers to questions 2 and 4 to calculate the reaction rate in moles per liter of acetone consumed per second (M/s).

6. Many reaction rates *roughly* double for each ten degree rise in temperature. Based on your results and those of your classmates, does the reaction in this experiment fit the "roughly double" pattern? Use data to support your answer.



Experiment 21

Testing Vinegar Concentration

Vinegar is a water solution of acetic acid. Federal regulations specify that any white vinegar product sold in the United States must be 5% acetic acid by mass. To ensure that such requirements are met, periodic routine analyses are run. Titration is a common technique for such analyses. A solution of known concentration is used to determine the concentration of another solution. Starting with a measured volume of one solution, you add the other solution (the *titrant*) at a controlled rate until reaction is complete. An indicator is used to determine the equivalence point of the reaction. When the indicator used in this experiment, phenolphthalein, turns from colorless to pale pink, your titration is done.

In this experiment, both the volume and the molarity of the titrant are known. Thus the number of moles of solute in the titrant can be calculated. With this data and the equation for the reaction, you can determine the number of moles of solute in the sample being analyzed. The equation for the reaction is shown below. Notice that one mole of acetic acid will neutralize one mole of sodium hydroxide.

$$HC_2H_3O_2(aq) + NaOH(aq) \longrightarrow NaC_2H_3O_2(aq) + H_2O(l)$$

You will prepare a standard sodium hydroxide solution, then use that solution to determine the percent of acetic acid by mass in white vinegar. You will determine the quantities used in each of four titrations by finding the mass of the pipets before and after titration.

OBJECTIVES

1. ***Prepare*** a sodium hydroxide solution of known molarity.
2. ***Titrate*** samples of white vinegar with the sodium hydroxide solution.
3. ***Determine*** the moles of sodium hydroxide used in each titration, and ***calculate*** the moles of acetic acid in each vinegar sample.
4. ***Determine*** the percent of acetic acid by mass in the vinegar.

Materials

Apparatus

centigram or milligram balance
50-mL beaker
25-mL graduated cylinder
4 10-mL Erlenmeyer flasks
forceps
2 5-mm microtip pipets
wash bottle with distilled water
safety goggles
laboratory apron
gloves

Reagents

1 g sodium hydroxide pellets
phenolphthalein solution in a microtip pipet
30 mL white vinegar

Caution: NaOH is corrosive. Wear gloves and safety goggles, and avoid all contact with skin and clothing. If any chemicals get into your eyes, use the eyewash fountain immediately. Wash spills or splashes off skin or clothing with a continuous stream of cool tap water and notify your teacher.

▶ Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–4 on the Report Sheet.

▶ Procedure

▶ NaOH will react rapidly with water vapor in the air. Because absorption of moisture will cause the pellets to stick together and the mass to change, step 2 must be completed carefully and quickly.

Part 1: Preparation of Standard NaOH

1. Put on your laboratory apron, safety goggles, and gloves.
2. Find the mass of a clean, dry 50-mL beaker. Using forceps, add enough NaOH pellets to give a mass of about 1.0 g. (Your teacher will tell you about how many to use.) Close the container *tightly*. Find the mass of the beaker and pellets. Record both masses in Table 1.
3. Add about 5–6 mL of distilled water to the beaker, then swirl to dissolve the pellets. Carefully feel the bottom of the beaker from time to time. Record your observations on the Report Sheet.
4. When the pellets have dissolved, pour the NaOH solution into a clean 25-mL graduated cylinder. (If the cylinder is not dry, rinse it with distilled water and drain well.) Use small amounts of distilled water from your wash bottle to rinse the beaker and add the rinsings to the cylinder. Continue until the total volume of the solution is precisely 25.0 mL. If you accidently exceed the 25.0-mL mark, dispose of the NaOH solution as described in Cleaning Up and repeat steps 2–4.
5. Wipe your beaker dry and then return the NaOH solution to it.

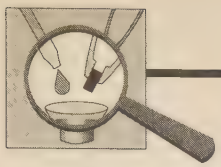
Part 2: Titration of Commercial Vinegar

6. Label one pipet *NaOH*; then fill it from the solution in your beaker. Fill another clean, dry pipet (labeled *vinegar*, or *VIN*) with vinegar. Determine the mass of each pipet and record it in Table 2.
7. Transfer about one-third of the vinegar solution into one of your flasks; then find the mass of the pipet. The mass of vinegar should be between 1 and 2 g; if it is not, increase the amount of vinegar for the other three titrations.
8. Add 2 drops of phenolphthalein indicator from the pipet labeled *PHTH*. While swirling, add the NaOH solution a few drops at a time until you get a pink color that does not fade with mixing and that lasts at least 20 seconds. The paler the pink, the better. Then find the mass of the NaOH pipet.
9. Repeat the titration three times, recording the mass of each pipet before and after each trial. You will probably need to refill one or both pipets from time to time. Use a clean flask for each titration. If necessary, wash the flasks between trials, rinsing with distilled water.

Part 3: Standardization of NaOH (optional)

Your teacher will provide an acid solution of known concentration, potassium acid phthalate (KHP). Run three titrations of KHP(aq) with NaOH(aq) to determine the precise concentration of the NaOH solution. Then recalculate the percent of acetic acid in vinegar.

▶ ***Cleaning Up*** The contents of your titration vessels may be rinsed down the drain with water. However, the unused NaOH solution should be neutralized. Add one drop of phenolphthalein to the NaOH beaker; then add vinegar until the pink color of the phenolphthalein disappears. The color change signals that neutralization is complete. Before leaving, clean up all materials and wash your hands thoroughly.



Experiment 21 Report Sheet

Testing Vinegar Concentration

Name _____

Class _____

Date _____

Prelab Questions

1. How does an indicator tell you when a titration is complete?

_____

_____

2. Why is it reasonable to assume that the mass in grams of the NaOH solution is equal to its volume in milliliters?

_____

_____

3. When working with a graduated cylinder, why shouldn't you fill it to above the top mark?

_____

_____

4. Briefly describe how you will dispose of the unused portion of the NaOH solution that you will prepare.

_____

_____

Data and Observations

Table 1

Observations (Part 1, step 3)

Mass of beaker and NaOH pellets:
Mass of empty beaker:
Mass of NaOH pellets:

Table 2

	Titration 1	Titration 2	Titration 3	Titration 4
Mass of vinegar pipet before titration				
Mass of vinegar pipet after titration				
Mass of vinegar				
Mass of NaOH pipet before titration				
Mass of NaOH pipet after titration				
Mass of NaOH				

Calculations

Show all calculations in the space provided.

1. Use the data from Part 1 to calculate the molarity of your NaOH solution.

Concentration of NaOH solution _____ M

2. Use the data in Part 2 to calculate the number of moles of NaOH used in each titration. (Because the density of vinegar is close to 1.0 g/mL, you may assume that the mass in grams of the solution is numerically equal to its volume in milliliters.)

Table 3

	Titration 1	Titration 2	Titration 3	Titration 4
Volume of NaOH used				
Moles of NaOH used				

3. Determine the number of moles of acetic acid present in each of your titration samples. Convert moles to grams. To find the percent of acetic acid in each vinegar sample, divide the mass of acetic acid by the mass of vinegar used, and convert the decimal fraction to a percent. Summarize your results in Table 4.

Table 4

	Titration 1	Titration 2	Titration 3	Titration 4
Moles NaOH				
Moles $HC_2H_3O_2$				
Mass of $HC_2H_3O_2$				
Mass Vinegar (g)				
% $HC_2H_3O_2$				

4. Calculate an average value for the percent of acetic acid in the vinegar you analyzed. Base your average on the three trials that show the closest agreement.

Percent of acetic acid in vinegar _____ %

5. Using your data from Tables 2 and 4, calculate the molar concentration of acetic acid in vinegar. (Assume the density of the vinegar is 1.0 g/mL.)

▶ Analysis and Conclusions

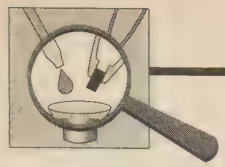
1. As noted in the introduction, vinegar should be 5% acetic acid by mass. Based on your results, does the vinegar you tested comply with the standard?

2. What could be done to improve the procedure to make the results more precise? Consider, for example, any assumptions that were made.

3. How many significant figures are you entitled to in your reported percent of acetic acid? Explain your answer.

4. How many significant figures does the manufacturer of the vinegar use? What effect, if any, could the number of significant digits have on the preparation of commercial vinegar?

5. Write a brief summary comparing your goals for this experiment with what you actually learned. Discuss any parts of the procedure that may have influenced your precision and describe how your results may have been affected.



Experiment 22

Redox Reactions

Reactions that involve the transfer of electrons from one atom (or ion) to another are called **redox** reactions from the terms **red**uction (electron gain) and **ox**idation (electron loss). The atom or ion that gains electrons is called an *oxidizing agent* because it causes the oxidation of the other species. The atom or ion that loses electrons is called a *reducing agent* because it provides the electrons for the reduction of the other species.

OBJECTIVES

1. ***Observe*** evidence of reaction when pairs of reactants are combined.
2. ***Determine*** which half-reactions occurred based on your observations.
3. ***Infer*** the relative reducing abilities of different reducing agents.

▶ Materials

Apparatus

96-well test plate
wash bottle
24-well plate (to hold reagent pipets)
forceps or tweezers
toothpicks as stirrers
safety goggles
laboratory apron
gloves

Reagents

1.0 M HCl(aq)
1.0 M $Cu(NO_3)_2$(aq)
1.0 M $Zn(NO_3)_2$(aq)
0.10 M $KMnO_4$(aq)
0.10 M $FeCl_3$(aq)
0.10 M $FeSO_4$(aq)
0.10 M NH_4VO_3(aq)
0.10 M $VOSO_4$(aq)
3% H_2O_2
0.10 M $CuSO_4$(aq)
5% NaOCl (bleach)
0.10 M Na_2SO_4(aq)
magnesium metal
copper metal
zinc metal

▶ Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–3 on the Report Sheet.

▶ Procedure

Part 1: Observing Redox Reactions

1. Put on your laboratory apron, safety goggles, and gloves.
2. Place 10 drops of 1.0 M HCl in a well of your test plate. Use forceps or tweezers to add a small piece of magnesium to the acid. Record your observations on the Report Sheet.

Caution: If you spill HCl on skin or clothing, rinse with a continuous stream of cool tap water.

3. Place 10 drops of 1.0 M copper(II) nitrate, $Cu(NO_3)_2$, in a well of your test plate. Add a small piece of zinc metal to the copper(II) nitrate solution. Record any evidence of reaction. (A change may not happen immediately.)
4. Place 10 drops of 1.0 M zinc nitrate, $Zn(NO_3)_2$, in another well of your test plate. Add a small piece of copper metal to the zinc nitrate and record any evidence of reaction.

Caution: $KMnO_4$ can stain skin and clothing. If you spill any, rinse it off in a continuous stream of cool tap water.

5. Place 2 drops of 0.1 M potassium permanganate solution, $KMnO_4$, in each of two wells of your test plate. Add 2 drops of 0.1 M iron(III) chloride (a source of Fe^{3+} ion) to one of the two wells containing the purple permanganate solution. To the other well, add 2 drops of 0.1 M iron(II) sulfate (a source of Fe^{2+} ion). A change will occur in one well. If no change is obvious, continue adding the appropriate iron solution to each well, a drop at a time, until a change *is* apparent. Note which iron ion reacts and record observed changes.
6. To the well in which a reaction occurred, continue adding the appropriate iron ion, drop by drop, until the color no longer changes. Record the number of drops that you used.

Part 2: Comparing Reducing Agents

▶ A change in the color of an aqueous solution of vanadium compounds is an indicator of a change in vanadium's oxidation state, which makes it easy to tell when an oxidation or reduction occurs.

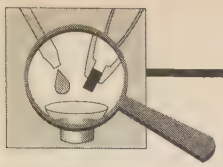
7. Observe and record the color of the 0.50 M ammonium vanadate solution, $NH_4VO_3(aq)$. The color is due to the VO_2^+ ion, which contains vanadium in the 5+ oxidation state. Observe and record the appearance of the 0.50 M vanadyl sulfate solution, $VOSO_4(aq)$. The color is due to the vanadyl ion, VO^{2+}, which contains vanadium in the 4+ state.
8. Select a clean row of wells on your test plate. Put 4 drops of the vanadyl sulfate solution in each of 8 wells. One drop at a time, add 4 drops of each test solution to the designated well.

Well	Test Solution	Species Being Tested
2	$KMnO_4$	MnO_4^-
3	$FeCl_3$	Fe^{3+}
4	bleach	OCl^-
5	Na_2SO_4	SO_4^{2-}
6	$CuSO_4$	Cu^{2+}
7	3% H_2O_2	H_2O_2

After each addition, record any changes that occur in Table 1 on the Report Sheet. (Use the end wells for comparison.) Look for a color *change,* not a change in the intensity of the existing color or a blending of colors such as the green produced when blue and yellow combine. If you do not observe a color change, write N.R. If you are uncertain, put 4 drops of the test solution either into an empty well or into a well that contains 4 drops of distilled water. If you are still uncertain, try adding the vanadyl solution to the test solution.

▶ ***Cleaning Up*** Use a wash bottle to rinse the solutions and any unreacted metal from your test plate onto paper towels in the container labeled *Experiment 22 Waste.* Return pipets and unused solutions to their proper place. Do not empty them into the sink.

9. Repeat step 8 using 4 drops of 0.10 M iron(II) sulfate solution, $FeSO_4(aq)$, in place of $VOSO_4(aq)$. Replace the $FeCl_3$ test solution with ammonium vanadate, $NH_4VO_3(aq)$, (a source of VO_2^+). Record any changes in Table 2.



Experiment 22 Report Sheet

Redox Reactions

Name _____

Class _____

Date _____

Prelab Questions

1. a. What are the most common ionic charges for the following elements?

Aluminum _____; Oxygen _____

b. Write half reactions showing aluminum and oxygen gas forming the ions you identified in 1a.

_____

c. Identify each of your reaction equations as being either an oxidation or a reduction.

_____

2. Give directions for preparing 100. mL of 0.10 M iron(III) chloride, $FeCl_3 \cdot 7H_2O$

_____

_____

3. If a household antiseptic solution is 3.0% H_2O_2 by mass, what is its molarity? Assume the solution has a density of 1.0 g/mL.

Data and Observations

Part 1: Observing Redox Reactions

Describe the reaction between magnesium and hydrochloric acid.

Describe the reaction between

a. copper(II) nitrate and zinc.

b. zinc nitrate and copper.

Describe the effect on $KMnO_4$ solution of

a. iron(II) sulfate solution.

b. iron(III) chloride solution.

For the iron solution that changed the color of the permanganate solution, how many drops were needed to complete the change?

_____ drops

Part 2: Comparing Reducing Agents

Table 1: Reactions with VO^{2+}

MnO_4^-	
Fe^{3+}	
OCl^-	
SO_4^{2+}	
Cu^{2+}	
H_2O_2	

Table 2: Reactions with Fe^{2+}

MnO_4^-	
VO_2^+	
OCl^-	
SO_4^{2+}	
Cu^{2+}	
H_2O_2	

Analysis and Conclusions

The list below contains six pairs of half-reactions. Each pair includes a reduction and its corresponding oxidation. Refer to this list as you answer the questions for Part 1.

Reductions	**Oxidations**
$Zn^{2+} + 2e^- \longrightarrow Zn$	$Zn \longrightarrow Zn^{2+} + 2e^-$
$Cu^{2+} + 2e^- \longrightarrow Cu$	$Cu \longrightarrow Cu^{2+} + 2e^-$
$Mg^{2+} + 2e^- \longrightarrow Mg$	$Mg \longrightarrow Mg^{2+} + 2e^-$
$Fe^{3+} + 1e^- \longrightarrow Fe^{2+}$	$Fe^{2+} \longrightarrow Fe^{3+} + 1e^-$
$2H^+ + 2e^- \longrightarrow H_2$	$H_2 \longrightarrow 2H^+ + 2e^-$
$Cl_2 + 2e^- \longrightarrow 2Cl^-$	$2Cl^- \longrightarrow Cl_2 + 2e^-$

Part 1: Observing Redox Reactions

1. Magnesium has two oxidation states, 0 for the neutral atom and 2+ for the magnesium ion.

   a. With which state did you begin in step 2?

   b. What was the state of magnesium after it reacted with HCl?

   c. Was magnesium oxidized or reduced? Explain.

   _____

   d. Write the half-reaction showing the change for magnesium.

   _____

2. The hydrochloric acid solution contains H^+ ions and Cl^- ions. Consult the list of half-reactions to determine which gas, H_2 or Cl_2, was a product of the reaction with Mg. (Remember that a complete reaction will involve an oxidation and a reduction.) Write the equation for the half-reaction that produced the gas.

   _____

   _____

3. Combine your half-reactions from questions 1 and 2 to write the net equation for the complete reaction.

   _____

4. Circle the equation that represents the reaction that occurred between zinc and copper and justify your choice.
   $Cu + Zn^{2+} \longrightarrow Cu^{2+} + Zn$ or $Cu^{2+} + Zn \longrightarrow Cu + Zn^{2+}$

   _____

5. For the equation you selected in question 4, identify each of the following: the species that was reduced, the species that was oxidized, the reducing agent, and the oxidizing agent.

   _____

6. Explain why it took more than two drops of the iron solution to react with the two drops of permanganate solution even though both solutions were 0.1 M.

   _____

   _____

7. **a.** In the reaction with MnO_4^-, one species of iron ion is converted to another. Circle the equation that represents the reaction that occurs.
   $Fe^{2+} \longrightarrow Fe^{3+} + e^-$ or $Fe^{3+} + e^- \longrightarrow Fe^{2+}$

   b. Is iron being oxidized or reduced? Explain.

   _____

   c. Is the permanganate ion being oxidized or reduced? Explain.

   _____

   d. Identify the oxidizing agent and the reducing agent in the reaction between permanganate and iron ions.

   _____

Part 2: Comparing Reducing Agents

8. Which test ions were able to oxidize VO^{2+} to VO_2^+.

9. Which test ions were able to change iron(II) to iron(III)?

10. Why was VO_2^+ substituted for Fe^{3+} as a test ion in step 9?

11. Which is more easily oxidized, Fe^{2+} or VO_2^+? Defend your choice.

12. (Optional) Write balanced equations for the reactions you observed in Part 2. Refer to the list of half-reactions below and these guidelines.

i. Include one oxidation and one reduction in each reaction system.

ii. Reduction half-reactions can be reversed to produce oxidation half-reactions.

iii. The number of electrons lost during oxidation must equal the number gained during reduction. You will have to multiply some half reactions by a factor to balance the numbers of electrons gained and lost. (Note: In strongly acidic solutions, the hypochlorite ion, OCl^-, is converted to hypochlorous acid, HOCl.)

$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$

$HOCl + H^+ + 2e^- \longrightarrow Cl^- + H_2O$

$Fe^{3+} + e^- \longrightarrow Fe^{2+}$

$SO_4^{2-} + 2H^+ + 2e^- \longrightarrow SO_3^{2-} + H_2O$

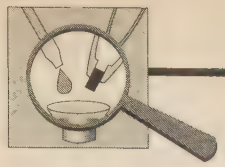
$H_2O_2 + 2H^+ + 2e^- \longrightarrow 2H_2O$

$Cu^{2+} + 2e^- \longrightarrow Cu$

$VO_2^+ + 2H^+ + e^- \longrightarrow VO^{2+} + H_2O$

Reactions from step 9:

Reactions from step 10:



Experiment 23

Testing for Water Hardness

Tap water left to boil for a long time deposits a white film on the walls of its container. You may have seen such a deposit on laboratory glassware. These deposits usually contain $CaCO_3$ and $CaSO_4$, as well as small amounts of $MgCO_3$ and $MgSO_4$. Water that contains Ca^{2+} and Mg^{2+} ions is called *hard water;* the greater the concentration, the harder the water. Soaps react with Ca^{2+} and Mg^{2+} to form a precipitate that is responsible for the gray film on clothes and "bath tub ring."

You will determine the degree of hardness of tap water. Because the ions are present in very small concentrations, you will measure them in parts per million (ppm). A calcium ion concentration, $[Ca^{2+}]$, of 100 parts per million represents 100 grams of Ca^{2+} in one million grams of water. There is generally far more Ca^{2+} than Mg^{2+} in water samples; thus water hardness tests is ordinarily reported in terms of Ca^{2+} concentration.

Ca^{2+} and Mg^{2+} ions have a strong attraction for a large molecule called ethylenediaminetetraacetic acid, EDTA. EDTA molecules surround and trap metal cations. The complex ion that forms is called a *chelate,* from the Greek word *chēlē,* meaning claw. The process is called *chelation.*

You will titrate water samples with a solution of EDTA until chelation is complete. To identify the end point, you will use an indicator called *calmagite,* which can also trap calcium and magnesium ions, but not as effectively as EDTA. Calmagite is red in the presence of Ca^{2+} and Mg^{2+} and blue in their absence. During your titrations, the indicator will turn from red to blue as EDTA takes the trapped metal ions away from the calmagite. If the solution is not moderately basic (pH 10), the end point may be impossible to detect. You will add a buffer to ensure that the system will be basic and that the end point is sharp.

OBJECTIVES

1. ***Titrate*** three tap water samples and one distilled water control.
2. ***Calibrate*** the drop size of the EDTA pipet used in the titrations.
3. ***Calculate*** the Ca^{2+} concentration for each sample and an average value for the water's hardness.

▶ Materials

Apparatus

10-mL graduated cylinder
4 50-mL beakers
thin-stem pipet

Reagents

tap water
distilled water
0.010 M EDTA solution in a microtip pipet
pH 10 buffer solution
calmagite indicator in thin-stem pipet
safety goggles

Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1 and 2 on the Report Sheet.

Procedure

Part 1: Titration Data

1. Put on your safety goggles before you begin.
2. Label four 50-ml beakers, A–D. Rinse a 10-mL graduate with tap water. Drain excess water from the cylinder, but do not dry the inside. Fill the cylinder to precisely 5.00 mL with tap water. Pour the water into Beaker A. Let the water drain from the cylinder to ensure as complete a transfer as possible. Repeat to fill Beakers B and C.
3. Rinse your graduate thoroughly with distilled water and drain as before to remove excess drops. Put precisely 5.00 mL of distilled water in your control beaker, D.
4. Add 2.0 mL of the pH 10 buffer, and 3 drops of calmagite indicator to each beaker, A–D. Swirl gently to mix.
5. Use a microtip pipet to add 1 drop of 0.010 M EDTA solution to beaker D and swirl to mix. If a clear blue color does not appear, continue adding drops of EDTA, swirling after each drop, until it does appear. Record the number of drops needed in Table 1. Set the control beaker aside to compare with your titration samples.
6. Add 10 drops of EDTA to beaker A and swirl to mix. If the solution does not turn blue, add 5 more drops of EDTA and swirl again. Continue in 5-drop increments—keeping track of the number of drops used—until you get a blue color that does not change back to red. You want a pure blue color with no trace of red or purple. (With very hard water, you might need more than 100 drops.) Record the total number of drops used in Table 1.
7. For samples B and C, follow the procedure in step 6 EXCEPT as you near the expected end point, add only a drop or two at a time, swirling the beaker after each addition. This modification should produce a more precise value. If your results from samples B and C vary greatly, do one more trial.

▶ If necessary, use a clean, thin-stem pipet to adjust the water level.

▶ Use the same EDTA pipet throughout the experiment. If necessary, refill it, but do not substitute a different pipet.

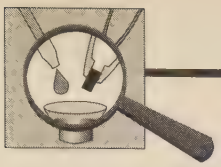
▶ When you compare beaker A with the control beaker, remember that the intensity of color may vary due to the difference in volume.

Part 2: Calibration of the EDTA Pipet

To convert your data from the number of drops of EDTA used to volumes in mL, you need to calibrate your EDTA pipet after you complete Part 1.

8. Add about 2 mL of water to your 10-ml graduated cylinder. Read the volume *to the nearest 0.01 mL* by estimating between scale markings and record it in Table 2.
9. Use your EDTA pipet to add exactly 30 drops of liquid to the water in the cylinder. Direct the drops toward the water surface; if they hit the sides, they may not reach the water in the bottom. Record the new volume to the nearest 0.01 mL.
10. Add a second 30 drops to the graduate and record the volume.

▶ ***Cleaning Up*** Rinse solutions down the drain with water. Wash glassware with soap and water, rinse, and leave to drain. Clean up your work area and wash your hands thoroughly before leaving the laboratory.



Experiment 23 Report Sheet

Testing for Water Hardness

Name _____

Class _____

Date _____

Prelab Questions

1. In Chapter 12, you learned that hard water contains dissolved calcium bicarbonate, $Ca(HCO_3)_2$, and magnesium bicarbonate, $Mg(HCO_3)_2$. However, the deposits left when hard water evaporates contain calcium carbonate, $CaCO_3$, and magnesium carbonate, $MgCO_3$.

a. What happens to the bicarbonate ions during evaporation?

_____

b. Write balanced equations for the reactions that take place when aqueous solutions of calcium bicarbonate and magnesium bicarbonate evaporate.

_____

_____

2. Hard water often has a $[Ca^{2+}]$ of 100 parts per million, which is equivalent to 100. mg/L.

a. Calculate the moles of calcium carbonate that would contain 100. mg of calcium ion.

b. Use the moles of calcium ion calculated in 2a to determine the molarity of a calcium carbonate solution with a $[Ca^{2+}]$ of 100. mg/L.

Data and Observations

Table 1: Titration Data

Drops of EDTA needed for control (beaker D):

Drops of EDTA needed for sample A:

Drops of EDTA needed for sample B:

Drops of EDTA needed for sample C:

Table 2: Calibration Data

Initial volume in graduated cylinder:

Volume with 30 drops EDTA added:

Volume with additional 30 drops EDTA:

▶ Calculations

Show all calculations for one of your titrations; report results only for the others.

1. Use your calibration data to determine the average volume of 30 drops, and convert the average to the number of drops equivalent to 1.00 mL.

Average number of drops per mL: _____

2. Determine the volume (in liters) of EDTA used for each trial from the number of drops you used in each titration. Show your calculations in the space below and enter your results in Table 3.

3. The concentration of the EDTA solution is 0.010 M (0.010 mol EDTA per liter of solution). Calculate the moles of EDTA used for each titration and add these results to Table 3.

4. One mole of EDTA combines with one mole of Ca^{2+}. For each trial, calculate (a) the moles of calcium ions that reacted and (b) the mass of these calcium ions. Add these results to Table 3.

5. If you assume the density of water to be 1.00 g/mL, then 1 liter of water contains 1000 grams, and 1 ppm of calcium would correspond to a mass of 0.001 g Ca^{2+} per liter. Use these facts to calculate the hardness of your samples in ppm as well as the average hardness for the water tested. Record your results in Table 3.

Table 3

	Sample A	Sample B	Sample C	Average
Volume of EDTA				
Moles of EDTA				
Moles of Ca^{2+}				
Mass of Ca^{2+}				
Mass of water				
$[Ca^{2+}]$ in ppm				

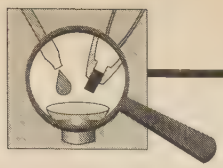
Analysis and Conclusions

1. Identify two main sources of experimental error in the procedure. If possible, suggest ways to modify the procedure that would minimize those errors.

2. (Optional.) The number of drops of EDTA needed to change the color of your distilled water control is called a *blank*. The usual technique in a titration is to subtract the volume of EDTA needed to change the color of the control from the EDTA volume used in each analysis. Subtract the blank for your control from your EDTA volume data and recalculate the hardness of your water. Discuss the results of this recalculation.

▶ Extensions (Optional)

1. If you have access to a commercial water softener, find out if it is possible to take 40–50 mL samples of the softened and unsoftened water without seriously disrupting the operation of the system. If so, obtain permission to analyze at least three samples of each. Discuss the significance of your results. If possible, try this analysis just before and just after the softener tank is recharged.
2. Analyze the filtered and the unfiltered water from a water supply that has at least one faucet with a purification filter and one without. Discuss your results.
3. Repeat Experiment 23, but use 2 mL of water and 1 mL of buffer, instead of 5 mL and 2 mL, respectively. Compare both the hardness values obtained and the consistency (precision) of data, and discuss your findings.
4. Repeat Experiment 23 with 10 mL of water and 4 mL of buffer. Compare both the hardness values obtained and the consistency (precision) of data, and discuss your findings.



Experiment 24

Separating Liquids by Distillation

As you learned in Chapter 14, the boiling point of a liquid is the temperature at which vapor pressure equals the external pressure on the liquid. If the liquid is in a closed container, some of the evaporated molecules strike the surface of the liquid and condense back into the liquid phase. A process called *distillation* can be used to keep the liquid that condenses from mixing with the original liquid.

Figure 1 on the next page shows the distillation apparatus you will use in this experiment. Vapor from the boiling liquid rises from the flask into the distilling column. The vapor that enters the side arm of the distilling column loses heat to the ice-water bath and condenses. The condensed liquid is called the *distillate.*

Distillation can be used to purify a liquid, separate liquids from solids, or separate a mixture of liquids. If the distillation flask held sea water, for example, the distillate would be pure water because the salt and other trace components are not volatile and would be left behind. Simple distillation works well when only one of the components of a mixture is volatile.

Miscible liquids can be separated by the difference in their boiling points. In this experiment, you will distill a 50:50 mixture of methanol and water. The boiling point of methanol is about 65°C; water boils at 100°C. You will collect the distillate that forms while the thermometer holds steady at about 65°C. You will stop collecting distillate when the temperature rises to 70°C. Your distillate will contain mainly methanol. It will not be pure because some water will have evaporated along with the methanol.

Ideally, you would collect a series of distillates until you reached the boiling point of water. Each distillate would correspond to a different temperature range. The distillates collected at the beginning would contain mainly methanol; those collected at the end of the process would contain mainly water. By further distillation of each distillate, you could eventually separate all the methanol from the water. Such a process is called *fractional distillation.* It is the process used to separate petroleum into its components.

OBJECTIVES

1. ***Distill*** a 50:50 mixture of methanol and water.
2. ***Graph*** the distillation data.
3. ***Compare*** the physical properties of the distillate with those of the mixture and the residue in the flask.

Materials

Apparatus

milligram or centigram balance
5 mL long-neck flask
distilling column with side arm
thermometer and adapter
connector with support rod
2 ring stands with clamps
sand bath
boiling stone
vial
30- or 50-mL beaker with chilled water or crushed ice
metal spatula or wood splint
3–5 mL calibrated syringe (optional)
safety goggles

Reagents

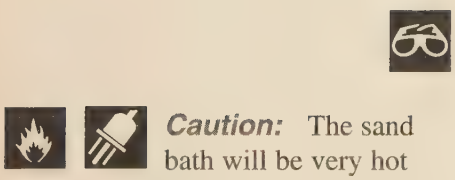
50:50 methanol–water mixture

Prelab

1. Read the introduction and procedure before you begin.
2. Answer Prelab questions 1–4 on the Report Sheet.

Procedure

Part 1: Assembling the Distillation Apparatus



Caution: The sand bath will be very hot (> 175°C) but will not appear so.

1. Put on your safety goggles before your begin.
2. Check to be sure the sand bath is on and that the controller has been set to 30% power. The surface of the sand should be level and smooth.
3. Refer to Figure 1 as you assemble the apparatus. Your instructor has pre-assembled the thermometer and the distilling column. Attach one side of the connector and support rod to the distilling column.

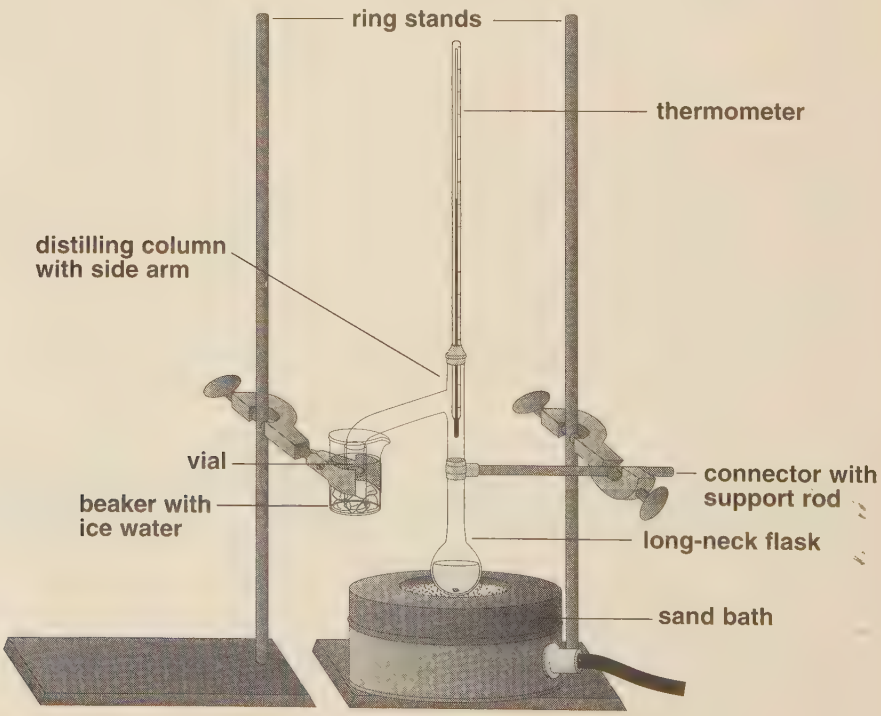


Figure 1: Distillation apparatus

4. Place a boiling stone in the long-neck flask. Then add 4.0 mL of the methanol–water mixture to the flask. Record the appearance of the mixture in Table 1 on the Report Sheet. Attach the flask to the connector and clamp the entire apparatus to the ring stand above the sand bath.
5. Clamp a small (30-mL or 50-mL) beaker of ice-water containing a vial to the second ring stand. The vial should fit snugly over the end of the side arm to minimize the escape of methanol vapors.
6. If necessary, clamp the thermometer so that the entire apparatus is vertical and the bulb of the thermometer does not touch the walls of the distilling column. Adjust the height of the clamps until the flask sits about 1 mm above the surface of the sand.

Caution: Methanol is very flammable. There should be no open flames in the lab during the procedure. Do not drop any of the mixture directly on the hot sand bath. If you get any methanol on your skin, wash with a stream of cool tap water.

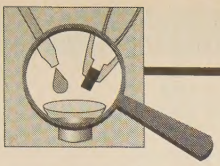
Part 2: The Distillation

7. If boiling does not begin within a minute or two, use a metal spatula or wood splint to pile up the sand around the bottom of the flask. If the boiling becomes too vigorous, slow it down by scraping sand away from the flask.
8. Record the temperature when bubbles first start to rise from the boiling stone in Table 2. Record the temperature every 30 seconds until the temperature reaches 70°C. Also record the temperature when a drop of liquid first appears on the bulb of the thermometer. Adjust the heat as described in step 7 so that the rate of condensation in the side arm *does not exceed 3 or 4 drops per minute.*
9. When the temperature exceeds 70°C, adjust the apparatus so that the flask is well above the sand. When liquid no longer collects in the vial, carefully lower the beaker and vial down from the side arm and adjust the apparatus so that the flask is no longer above the sand bath.
10. Remove the vial from the cold water, dry the outside, and record the appearance of the distillate.
11. (Optional) Determine the density of your distillate. Find the mass of a small dry container. Use a syringe to withdraw a portion of the distillate from the vial. Transfer a known volume to the container. Return the unused distillate to the vial. Then determine the mass of the container and distillate sample. Record the data in Table 3.
12. (Optional) If your teacher demonstrates the differences in flammability among the distillate, the residual liquid in the flask, and the original mixture, record your observations of the demonstration in Table 1.

Caution: Do not allow the liquid in the distilling flask to completely evaporate. A dry flask is likely to break when heated.

▶ Cleaning Up

The liquids can be recycled. After the flask and distilling column have cooled, *carefully* disconnect the flask from the column. Record the appearance of the liquid remaining in the flask and compare its odor to that of the distillate. (Use standard precautions for smelling a vapor.) Then pour the liquids into the appropriate containers. Discard the boiling stone in the wastebasket.



Experiment 24 Report Sheet

Separating Liquids by Distillation

Name _____

Class _____

Date _____

Prelab Questions

1. How will you control the rate of heating during the distillation?

_____

_____

2. Why is the thermometer bulb placed just below the side arm of the distilling column?

_____

_____

3. What is the function of the ice water in the beaker that holds the vial?

_____

_____

4. What is the purpose of the boiling stone?

_____

_____

_____

Data and Observations

Table 1: Comparison Of Properties

	Appearance	Flammability
Mixture		
Distillate		
Liquid residue		

Comparison of odor of distillate and residue:

Table 2: Temperature Readings

Temperature at first sign of boiling (time 0.0 min):

Temperature when drop of liquid first appears on thermometer bulb:

Time (min)	Temperature (°C)	Time (min)	Temperature (°C)	Time (min)	Temperature (°C)

▶ Calculations

Table 3: Density of Distillate (optional)

Mass of small, dry container:

Volume of distillate:

Mass of container and distillate:

Mass of distillate:

Density of distillate:

Complete Table 3. Calculate the mass of distillate collected. Calculate the density of the distillate.

▶ Analysis and Conclusions

1. Make a graph of the boiling temperature of the distillate versus time. Describe the shape of your graph.

2. At what temperature did you collect most of the distillate? Compare that temperature with the boiling point of pure methanol? Is it reasonable to assume that the distillate is, in fact, methanol? What can you infer about the liquid remaining in the flask? Explain.

3. What properties other than density can be used to verify the identity of an unknown liquid?

Periodic Table of the Elements

(based on $^{12}_{6}C = 12.0000$)

1* 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 **H** Hydrogen 1.008 $1s^1$																	2 **He** Helium 4.003 $1s^2$
2 **Li** Lithium 6.941 $2s^1$	4 **Be** Beryllium 9.012 $2s^2$											5 **B** Boron 10.811 $2s^22p^1$	6 **C** Carbon 12.011 $2s^22p^2$	7 **N** Nitrogen 14.007 $2s^22p^3$	8 **O** Oxygen 15.999 $2s^22p^4$	9 **F** Fluorine 18.998 $2s^22p^5$	10 **Ne** Neon 20.180 $2s^22p^6$
3 **Na** Sodium 22.990 $3s^1$	12 **Mg** Magnesium 24.305 $3s^2$											13 **Al** Aluminum 26.982 $3s^23p^1$	14 **Si** Silicon 28.086 $3s^23p^2$	15 **P** Phosphorus 30.974 $3s^23p^3$	16 **S** Sulfur 32.066 $3s^23p^4$	17 **Cl** Chlorine 35.453 $3s^23p^5$	18 **Ar** Argon 39.948 $3s^23p^6$
4 **K** Potassium 39.098 $4s^1$	20 **Ca** Calcium 40.078 $4s^2$	21 **Sc** Scandium 44.956 $4s^23d^1$	22 **Ti** Titanium 47.88 $4s^23d^2$	23 **V** Vanadium 50.942 $4s^23d^3$	24 **Cr** Chromium 51.996 $4s^13d^5$	25 **Mn** Manganese 54.938 $4s^23d^5$	26 **Fe** Iron 55.847 $4s^23d^6$	27 **Co** Cobalt 58.933 $4s^23d^7$	28 **Ni** Nickel 58.693 $4s^23d^8$	29 **Cu** Copper 63.546 $4s^13d^{10}$	30 **Zn** Zinc 65.39 $4s^23d^{10}$	31 **Ga** Gallium 69.723 $4s^24p^1$	32 **Ge** Germanium 72.61 $4s^24p^2$	33 **As** Arsenic 74.922 $4s^24p^3$	34 **Se** Selenium 78.96 $4s^24p^4$	35 **Br** Bromine 79.904 $4s^24p^5$	36 **Kr** Krypton 83.80 $4s^24p^6$
5 **Rb** Rubidium 85.468 $5s^1$	38 **Sr** Strontium 87.62 $5s^2$	39 **Y** Yttrium 88.906 $5s^24d^1$	40 **Zr** Zirconium 91.224 $5s^24d^2$	41 **Nb** Niobium 92.906 $5s^14d^4$	42 **Mo** Molybdenum 95.94 $5s^14d^5$	43 **Tc** Technetium (98) $5s^24d^5$	44 **Ru** Ruthenium 101.07 $5s^14d^7$	45 **Rh** Rhodium 102.906 $5s^14d^8$	46 **Pd** Palladium 106.42 $4d^{10}$	47 **Ag** Silver 107.868 $5s^14d^{10}$	48 **Cd** Cadmium 112.411 $5s^24d^{10}$	49 **In** Indium 114.82 $5s^25p^1$	50 **Sn** Tin 118.710 $5s^25p^2$	51 **Sb** Antimony 121.757 $5s^25p^3$	52 **Te** Tellurium 127.60 $5s^25p^4$	53 **I** Iodine 126.904 $5s^25p^5$	54 **Xe** Xenon 131.29 $5s^25p^6$
6 **Cs** Cesium 132.905 $6s^1$	56 **Ba** Barium 137.327 $6s^2$	57 **La** Lanthanum 138.906 $6s^25d^1$	72 **Hf** Hafnium 178.49 $6s^25d^2$	73 **Ta** Tantalum 180.948 $6s^25d^3$	74 **W** Tungsten 183.85 $6s^25d^4$	75 **Re** Rhenium 186.207 $6s^25d^5$	76 **Os** Osmium 190.2 $6s^25d^6$	77 **Ir** Iridium 192.22 $6s^25d^7$	78 **Pt** Platinum 195.08 $6s^15d^9$	79 **Au** Gold 196.967 $6s^15d^{10}$	80 **Hg** Mercury 200.59 $6s^25d^{10}$	81 **Tl** Thallium 204.383 $6s^26p^1$	82 **Pb** Lead 207.2 $6s^26p^2$	83 **Bi** Bismuth 208.980 $6s^26p^3$	84 **Po** Polonium (209) $6s^26p^4$	85 **At** Astatine (210) $6s^26p^5$	86 **Rn** Radon (222) $6s^26p^6$
7 **Fr** Francium (223) $7s^1$	88 **Ra** Radium (226.025) $7s^2$	89 **Ac** Actinium 227.028 $7s^26d^1$	104† **Rf** Rutherfordium (261) $7s^26d^2$	105† **Ha** Hahnium (262) $7s^26d^3$	106† **Sg** Seaborgium (263) $7s^26d^4$	107† **Ns** Neilsbohrium (262) $7s^26d^5$	108† **Hs** Hassium (265) $7s^26d^6$	109 **Mt** Meitnerium (266) $7s^26d^7$	110								

14 — Atomic number
Si — Symbol
Silicon — Name
28.086 — Atomic mass
$3s^23p^2$ — Valence electron configuration

TRANSITION METALS

An atomic mass given in parentheses is the mass number of the isotope of longest half-life for that element.

**The 1–18 group designation has been recommended by the International Union of Pure and Applied Chemistry (IUPAC) but is not widely used. In this book, we refer to the standard U.S. notation for group numbers (1A–8A and 1B–8B).*

†ACS-endorsed names; IUPAC names for 104–108 are dubnium (Db), joliotium (Jl), rutherfordium (Rf), bohrium (Bh), and hahnium (Hn), respectively.

INNER TRANSITION METALS

Lanthanide series	58 **Ce** Cerium 140.115 $6s^24f^15d^1$	59 **Pr** Praseodymium 140.908 $6s^24f^3$	60 **Nd** Neodymium 144.24 $6s^24f^4$	61 **Pm** Promethium (145) $6s^24f^5$	62 **Sm** Samarium 150.36 $6s^24f^6$	63 **Eu** Europium 151.965 $6s^24f^7$	64 **Gd** Gadolinium 157.25 $6s^24f^75d^1$	65 **Tb** Terbium 158.925 $6s^24f^9$	66 **Dy** Dysprosium 162.50 $6s^24f^{10}$	67 **Ho** Holmium 164.930 $6s^24f^{11}$	68 **Er** Erbium 167.26 $6s^24f^{12}$	69 **Tm** Thulium 168.934 $6s^24f^{13}$	70 **Yb** Ytterbium 173.04 $6s^24f^{14}$	71 **Lu** Lutetium 174.967 $6s^24f^{14}5d^1$
Actinide series	90 **Th** Thorium 232.038 $7s^26d^2$	91 **Pa** Protactinium 231.036 $7s^25f^26d^1$	92 **U** Uranium 238.029 $7s^25f^36d^1$	93 **Np** Neptunium 237.048 $7s^25f^46d^1$	94 **Pu** Plutonium (244) $7s^25f^6$	95 **Am** Americium (243) $7s^25f^7$	96 **Cm** Curium (247) $7s^25f^76d^1$	97 **Bk** Berkelium (247) $7s^25f^9$	98 **Cf** Californium (251) $7s^25f^{10}$	99 **Es** Einsteinium (252) $7s^25f^{11}$	100 **Fm** Fermium (257) $7s^25f^{12}$	101 **Md** Mendelevium (258) $7s^25f^{13}$	102 **No** Nobelium (259) $7s^25f^{14}$	103 **Lr** Lawrencium (260) $7s^25f^{14}6d^1$

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